CHAPTER 2 | Atoms, Ions, and Molecules: The Building Blocks of Matter

2.1. **Collect and Organize**
This question asks us to identify the element in Figure P2.1 that has the fewest protons in its nucleus.

Analyze
An element is defined by the number of protons in its nucleus; as we move to higher elements in the periodic table, the number of protons increases.

Solve
The lightest, lowest atomic number element highlighted is hydrogen (purple). Hydrogen therefore has the fewest protons (one) in its nucleus.

Think About It
The other elements highlighted have more protons in their nucleus: helium (blue) has 2, fluorine (green) has 9, sulfur (yellow) has 16, and arsenic (red) has 33.

2.2. **Collect and Organize**
This question asks us to identify the element in Figure P2.1 that has the most neutrons in its nucleus.

Analyze
As we move to higher atomic numbers in the periodic table, the number of neutrons in the nucleus increases.

Solve
The largest atomic number element highlighted is arsenic. Arsenic (red), therefore, has the most neutrons in its nucleus.

Think About It
Arsenic is a monoisotopic element; its only stable isotope is ^{75}As with 42 neutrons in its nucleus along with 33 protons.

2.3. **Collect and Organize**
We are asked to identify which shaded element in Figure P2.1 is stable and yet has no neutrons in its nucleus.

Analyze
The presence of neutrons in nuclei helps to overcome the repulsive forces of more than one proton in the nucleus.

Solve
The only shaded element that does not have more than one proton in its nucleus and therefore would be stable without neutrons is the element shaded purple, hydrogen.

Think About It
Without neutrons, nuclei with more than one proton would fly apart.

2.4. **Collect and Organize**
From the shaded elements in Figure P2.4, we are to choose the one for which no stable isotopes exist.

Analyze
When an element has no stable isotopes, the element is radioactive. Starting with bismuth (atomic number 83), all the elements are radioactive.

Solve
Lawrencium (shaded light orange) is the only shaded element with an atomic number greater than 83, so it is the unstable element in Figure P2.4.

Think About It

The last naturally occurring element found on Earth is uranium (element 92); the rest, through the newest discovered elements, oganesson (Z = 118) in 2002 and tennessine (Z = 115) in 2010, are synthetic.

2.5. **Collect and Organize**

From among the shaded elements in the periodic table shown in Figure P2.4, we are to identify a transition metal, an alkali metal, a halogen, and a chemically inert gas.

Analyze

Transition metals occupy the middle of the periodic table and reside in groups 3–12, alkali metals occupy the first column of the periodic table (group 1), and halogens occupy the second-to-last column on the right (group 17), next to the noble gases.

Solve

(a) The transition metal element is shaded green and is gold (Au).
(b) The alkali metal is shaded blue and is sodium (Na).
(c) The halogen is shaded lilac and is chlorine (Cl).
(d) The chemically inert gas is shaded red and is neon (Ne).

Think About It

A few other regions of the periodic table also have names. Can you identify the alkaline earth metals, the noble gases, and the chalcogens?

2.6. **Collect and Organize**

For each of the nuclides depicted, we are to write the nuclide symbols.

Analyze

Nuclide symbols ^{A}X are written to identify both the element X (from the number of protons in the nuclide, denoted as the atomic symbol) and the mass number A (the total number of protons and neutrons in the nuclide, denoted as a superscript to the upper left of the element symbol).

Solve

(a) This nuclide has 3 protons and 4 neutrons; it is Li with a mass number of 7: ^{7}Li.
(b) This nuclide has 3 protons and 3 neutrons; it is Li with a mass number of 6: ^{6}Li.
(c) This nuclide has 6 protons and 8 neutrons; it is C with a mass number of 14: ^{14}C.
(d) This nuclide has 7 protons and 7 neutrons; it is N with a mass number of 14: ^{14}N.

Think About It

Carbon-14 is an unstable isotope of carbon with a half-life of 5,730 years and is used in radiocarbon dating

2.7. **Collect and Organize**

Using Figure P2.7 and our knowledge of the charges and the masses of alpha (α) and beta (β) particles, we are to determine which arrow (red, blue, or green) represents each particle's behavior as it moves through an electric field.

Analyze

Alpha particles are much more massive than beta particles and they have a positive charge, whereas beta particles have a negative charge.

Solve

Alpha particles, with their positive charges, will be deflected toward the negative side of the electric field. Their behavior is represented by the red arrow. Beta particles, however, are deflected toward the positive side of the electric field because they carry a negative charge; they are represented by the green arrow. Notice, too, that the alpha particles travel farther and with less arc in their deflection than the lighter beta particles. Because alpha particles have greater mass, their momentum through the electric field is higher and it would take a stronger field to deflect them to the same degree.

Think About It

The blue arrow might represent either a neutral particle or electromagnetic radiation which are not deflected by an electric field.

2.8. **Collect and Organize**

For this question we are asked to identify the subatomic particle that would curve in the same direction as the red arrow in Figure P2.7.

Analyze

The red arrow shows that the particle is deflected toward the negative plate, which means that the subatomic particle carries a positive charge.

Solve

Protons with their positive charge would curve in the same direction as alpha particles, the red arrow in Figure P2.7.

Think About It

Another subatomic particle represented in the diagram are neutrons (blue arrow).

2.9. **Collect and Organize**

Using Figure P2.9, we are to determine whether the mass spectrum shown is for dichloromethane or cyclohexane.

Analyze

Dichloromethane has a molar mass of 84.93 g/mol and cyclohexane has a molar mass of 84.15 g/mol. Those masses are close to each other, so we will have to use other clues from the mass spectrum to help. The figure shows two obvious mass peaks at 86 and 88 u and another grouping around 49 u.

Solve

The molecular ion peaks showing up at 84, 86, and 88 u show that this compound has isotopes of significant abundance. Cyclohexane has only carbon and hydrogen atoms and the abundance of C and H isotopes are too low to show this significantly in the mass spectrum. The *m*/*z* peak at 84 u is due to the presence of two atoms of ^{35}Cl ($CH_2{}^{35}Cl_2$), the *m*/*z* peak at 86 u is due to the presence of one atom of ^{35}Cl and one atom of ^{37}Cl ($CH_2{}^{35}Cl^{37}Cl$), and the *m*/*z* peak at 88 u is due to the presence of two atoms of ^{37}Cl ($CH_2{}^{37}Cl_2$). Also, the *m*/*z* peak grouping around 50 u is consistent with dichloromethane, CH_2Cl_2, losing one ^{35}Cl or one ^{37}Cl atom.

Think About It

^{35}Cl is about 76% naturally abundant and ^{37}Cl is about 24% naturally abundant. The relative intensities of the isotopic peaks in the mass spectrum also reflect those differences in abundance.

2.10. **Collect and Organize**

Given that krypton has six stable isotopes, we are to determine from the mass spectrum in Figure P2.10 the number of neutrons in the most abundant isotope and write the symbol for that isotope.

Analyze

The mass spectrum shows *m*/*z* peaks for masses of 78, 80, 82, 83, 84, and 86 u, of which the peak at 84 u has the highest relative intensity.

Solve

The most abundant isotope has the highest relative intensity; for krypton that is an *m*/*z* of 84 u. Krypton has 36 protons in its nucleus, and with the neutron having about the same mass as the proton that means that the most abundant isotope has 84 – 36 = 48 neutrons. The isotope symbol is ^{84}Kr.

Think About It

The other isotopes have 42, 44, 46, 47, and 50 neutrons in their nuclei, respectively.

2.11. **Collect and Organize**

This question asks us to correlate the position of an element in the periodic table with typical charges on the ions for the groups (or families) of elements.

Analyze

Figure 2.11 in the textbook shows the common charges on the elements used in forming compounds. That figure will help us determine which elements in monatomic form give the charges named in the question.

Solve

Highlighted elements in Figure P2.11 are K, Mg, Sc, Ag, O, and I.

(a) Elements in group 1 form 1+ ions, so K will form K^+ (dark blue). Silver (green) also typically forms a 1+ cation.
(b) Elements in group 2 form 2+ ions, so Mg forms Mg^{2+} (gray).
(c) Elements in group 3 form 3+ ions, so Sc forms Sc^{3+} (yellow).
(d) Elements in group 17 (the halogens) form 1– ions, so I forms I^- (purple).
(e) Elements in group 16 form 2– ions, so O forms O^{2-} (red).

Think About It

Elements on the left-hand side of the periodic table form cations, and elements on the right-hand side tend to form anions.

2.12. **Collect and Organize**

Using the representations of molecules, atoms, and ions in Figure P2.12 and the atomic color palette, we are asked to match the representations to the descriptions.

Analyze

From the color palette we find what each color indicates: sulfur (yellow), oxygen (red), carbon (black), potassium (purple), iodine (lilac), neutron (dark gray), proton (white). We also note that [B], [D], [F], and [H] represent nuclei; [A] and [I] represent ionic solids; and [C], [E], and [G] represent molecules.

Solve

(a) Potassium iodide (KI) has a one-to-one ratio of K^+ to I^-; potassium oxide (K_2O) has a two-to-one ratio of K^+ to O^{2-}. Representation [A] has ions in equal ratios and is therefore KI, whereas representation [I] has fewer red ions than purple ions and is therefore K_2O.
(b) When we count the gray particles in representations [B], [D], [F], and [H], the order by increasing number of neutrons is [H] (5 neutrons) < [D] (6 neutrons) < [F] (7 neutrons) < [B] (8 neutrons).
(c) Isotopes have the same number of protons but a different number of neutrons. Representations [B] and [H] both have 6 protons but 8 and 5 neutrons, respectively.
(d) The molecule shown in [E] is CO_2, which has a mass of 12.011 + 2(15.999) = 44.009 g/mol.
(e) The molecule shown in [C] is SO_3 (molar mass = 80.066 g/mol and has one sulfur atom in the molecule), and the molecule shown in [G] is SO_2 (molar mass = 64.066 g/mol and also has one sulfur atom per molecule). Because SO_2 has a lower molar mass, 100 g of SO_2 (molecule [G]) would contain more molecules than 100 g of SO_3 (molecule [C]).
(f) The molecule shown in [C] is SO_3 (molar mass = 80.066 g/mol) and the molecule shown in [G] is SO_2 (molar mass = 64.066 g/mol). Because SO_2 has a lower molar mass, 100 g of SO_2 (molecule [G]) would contain more atoms of sulfur than 100 g of SO_3 (molecule [C]).

Think About It

You could also calculate the answers to e and f:

$$\text{(e) } 100 \text{ g SO}_3 \times \frac{1 \text{ mol}}{80.066 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 7.52 \times 10^{23} \text{ SO}_3 \text{ molecules}$$

$$100 \text{ g SO}_2 \times \frac{1 \text{ mol}}{64.066 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 9.40 \times 10^{23} \text{ SO}_3 \text{ molecules}$$

$$\text{(f) } 100 \text{ g SO}_3 \times \frac{1 \text{ mol}}{80.066 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \times \frac{1 \text{ S atom}}{\text{molecule}} = 7.52 \times 10^{23} \text{ S atoms in SO}_3$$

$$100 \text{ g SO}_2 \times \frac{1 \text{ mol}}{64.066 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \times \frac{1 \text{ S atom}}{\text{molecule}} = 9.40 \times 10^{23} \text{ S atoms in SO}_2$$

2.13. **Collect and Organize**
In this question we are asked to explain how Rutherford's gold-foil experiment changed the plum-pudding model of the atom to the nuclear model.

Analyze
The plum-pudding model of the atom viewed the electrons as small particles ("plums") in a diffuse, positively charged atom ("pudding"). In Rutherford's experiment, most of the alpha (α) particles (positively charged particles) directed at the gold foil went straight through, but a few of them bounced back toward the source of the α particles.

Solve
The plum-pudding model could not explain the infrequent large-angle deflections of alpha particles that Rutherford's students observed. However, these deflections could be explained by the particles' colliding with tiny, dense atomic nuclei that took up little of the volume of gold atoms, but that contained all the positive charge and most of the atoms' mass (Rutherford's model).

Think About It
The nucleus is about 10^{-15} m in diameter, whereas the atom is about 10^{-10} m. That size difference has often been compared to "a fly in a cathedral."

2.14. **Collect and Organize**
We are asked to consider what Rutherford's gold-foil experiments would have shown if the plum-pudding model had been valid.

Analyze
The plum-pudding model of the atom viewed the electrons as small particles in a diffuse, positively charged atom.

Solve
Had the plum-pudding model been valid, the α particles would have encountered only a diffuse positive charge when entering the atom, and they would have been only slightly deflected from their straight-line path by the electrons (Figure A2.14).

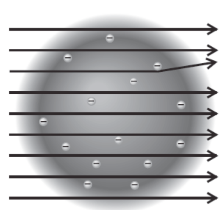


Figure A2.14

Think About It
Because the positive charge in the atom was viewed as diffuse, no α particles would have bounced back toward their source had the experiment confirmed the plum-pudding model. The key result of Rutherford's experiment

was that a few α particles came directly back toward the α source, indicating a highly dense nucleus of positive charge.

2.15. **Collect and Organize**

In this question we are to explain how J. J. Thomson discovered that cathode rays were not rays of pure energy but were actually charged particles.

Analyze

Thomson's experiment directed the cathode rays through a magnetic field, and he discovered that the ray was deflected. A magnetic field would not deflect energy (electromagnetic) rays.

Solve

When Thomson observed cathode rays being deflected by a magnetic field, he reasoned that the rays were streams of charged particles because only moving charged particles would interact with a magnetic field. Pure-energy rays would not.

Think About It

Thomson's discovery of the electron in cathode rays did not eliminate the use of the term *cathode ray.* CRTs (cathode-ray tubes) were the original screens (that is, not LCDs) used in televisions and computers.

2.16. **Collect and Organize**

We are to describe two differences between α and β particles.

Analyze

Both α and β particles are the result of nuclear processes, but they differ in their masses and charges.

Solve

Alpha particles are massive (about the mass of the helium nucleus), whereas β particles are very light (about the mass of an electron). These two particles also have opposite charges: α particles have a 2+ charge, whereas β particles have a 1– charge.

Think About It

Alpha particles are composed of two protons and two neutrons, and so when they are emitted from a nucleus the mass of the nuclide goes down by 4 atomic mass units and its atomic number is reduced by 2. Beta emission changes a neutron into a proton in the nucleus, and therefore the atomic number increases, but the mass of the nuclide changes very little.

2.17. **Collect and Organize**

Helium is found in pitchblende, an ore of radioactive uranium oxide found on Earth. We are asked to explain why helium is present in that ore.

Analyze

Pitchblende contains uranium oxide, and uranium is a naturally occurring radioactive element.

Solve

The helium is present because uranium (and some of the products of its nuclear reactions) undergoes α decay. The α particles, composed of two protons and two neutrons, easily pick up electrons from their environment to become helium atoms.

Think About It

All the helium on Earth is generated in this fashion and trapped. Helium, though, once in the atmosphere, escapes into space because it is so light.

2.18. **Collect and Organize**

In this thought experiment we are asked to predict how the result would have been affected if Rutherford had used a thicker piece of gold foil in his experiment.

Analyze

With thicker foil, the beam of α particles would encounter more gold atoms as they passed through the foil.

Solve

Fewer α particles would have passed directly through the foil and more would have been deflected at small and large angles. More α particles might also have been absorbed into the foil if thick enough, with less deflection seen.

Think About It

The thinness of the foil might have been crucial to seeing Rutherford's result.

2.19. **Collect and Organize**

In this thought experiment we are asked to predict what would have happened if Rutherford's gold atoms in the foil had absorbed β particles.

Analyze

If β particles were absorbed by the gold nuclei, that would mean that the β particles were somehow "reacting" with the nuclei of the gold atoms.

Solve

If a nucleus absorbs a β particle, a proton becomes a neutron. Therefore, we would observe the gold with 79 protons become platinum with 78 protons. Because a proton is converted into a neutron (and we assume we don't see any particle emission), the mass number of the gold and the platinum would be the same.

Think About It

The nuclear reaction that would describe the transmutation of gold-197 into platinum is

$$^{197}\text{Au} + \beta \rightarrow {}^{197}\text{Pt}$$

2.20. **Collect and Organize**

In this thought experiment we are asked to explain why aluminum and silver foils might have deflected fewer α particles than gold in the Rutherford experiment.

Analyze

Here we want to look at how gold differs from silver and aluminum in terms of their numbers of protons that can deflect α particles.

Solve

Silver and aluminum have smaller nuclei with fewer protons (47 and 13 protons, respectively) and so would not deflect like-charged α particles as well as gold would with 79 protons.

Think About It

Can you name a metal that might deflect more α particles than gold in the Rutherford experiment?

2.21. **Collect and Organize**

This question asks us to consider the ratio of neutrons to protons in an element where we are given the fact that the mass number is more than twice the atomic number.

Analyze

We can find the number of neutrons for an isotope by relating the number of protons to the mass number. From that result we can then determine the neutron-to-proton ratio.

Solve

We are given an isotope in which the mass number is more than twice the number of protons. With m being the mass number and p the number of protons, we can express this relationship as

$$m > 2p$$

The mass number is also equal to the number of protons plus the number of neutrons (n),

$$m = p + n$$

Combining these expressions

$$p + n > 2p$$

and solving for n gives

$$n > 2p - p$$
$$n > p$$

Therefore, the number of neutrons in this isotope is greater than the number of protons, and the neutron-to-proton ratio is greater than 1.

Think About It

We wouldn't have had to express the relationships between the nuclear particles mathematically if the isotope had a mass number equal to twice the number of protons. Then the number of neutrons would have to be the same as the number of protons, giving a neutron-to-proton ratio of 1:1.

2.22. **Collect and Organize**

For this question we are to compare the numbers of neutrons, protons, nucleons, and electrons in isotopes of oxygen.

Analyze

Atomic number gives the number of protons in the nucleus, and A, or the mass number, of the isotopes gives the number of *nucleons* (protons + neutrons in a nucleus). The number of neutrons can be calculated by subtracting the number of protons from the mass number. Finally, the number of electrons equals the number of protons in a neutral atom.

Solve

(a) The number of neutrons in each isotope is: ^{16}O, 8; ^{17}O, 9; and ^{18}O, 10.
(b) The number of protons in each isotope is: ^{16}O, 8; ^{17}O, 8; and ^{18}O, 8.
(c) The number of nucleons in each isotope is: ^{16}O, 16; ^{17}O, 17; and ^{18}O, 18.
(d) The number of electrons in each isotope is: ^{16}O, 8; ^{17}O, 8; and ^{18}O, 8.

Think About It

Isotopes are defined as elements with different numbers of neutrons in their nuclei, but having the same number of protons.

2.23. **Collect and Organize**

Given that most stable nuclides have at least as many neutrons in their nuclei as protons (and often more), we are to identify which element is an exception to that rule.

Analyze

Neutrons help stabilize the nucleus by counteracting the repulsive forces between protons.

Solve

Hydrogen, with only one proton, does not need neutrons to be stable and so ^{1}H is the exception.

Think About It

Hydrogen, however, can have one (for deuterium, ^{2}H) and even two (for tritium, ^{3}H) neutrons in its nucleus.

2.24. **Collect and Organize**

We are to explain why the mass number does not change when ^{19}F becomes $^{19}F^{-}$.

Analyze
The mass number, A, gives the sum of protons and neutrons in the nucleus. It does not include the mass of the electrons outside of the nucleus. Electrons do not appreciably add to the mass of the atom.

Solve
In adding an electron to ^{19}F to make the fluoride anion, we are not adding the electron to the nucleus, so that it would change the number or type of nucleons. Rather, an electron is added to the outside of the nucleus when an atom becomes a negative ion.

Think About It
Likewise, when an atom becomes a positive ion (cation) by losing an electron, the atomic number does not change.

2.25. **Collect and Organize**
For each element in this question, we must look at the relationship of the neutrons, protons, and electrons. We need to determine the element's atomic number from the periodic table and, from the mass number given, compute the number of neutrons to give the indicated isotope.

Analyze
An isotope is given by the symbol $^{A}_{Z}X$, where X is the element's symbol from the periodic table, Z is the atomic number (the number of protons in the nucleus), and A is the mass number (the number of protons and neutrons in the nucleus). Often, Z is omitted because the element symbol gives us the same information about the identity of the element. To determine the number of neutrons in the nucleus for each named isotope, we subtract Z (number of protons) from A (mass number). If the elements are neutral (no charge), the number of electrons equals the number of protons in the nucleus.

Solve

	Atom	Mass Number	Atomic Number = Number of Protons	Number of Neutrons = Mass Number – Atomic Number	Number of Electrons = Number of Protons
(a)	^{14}C	14	6	8	6
(b)	^{59}Fe	59	26	33	26
(c)	^{90}Sr	90	38	52	38
(d)	^{210}Pb	210	82	128	82

Think About It
Isotopes of an element contain the same number of protons but a different number of neutrons. Thus, isotopes have different masses.

2.26. **Collect and Organize**
For each element in this question, we must look at the relationship of the neutrons, protons, and electrons. We need to determine the element's atomic number from the periodic table and, from the mass number given, compute the number of neutrons to give the indicated isotope.

Analyze
An isotope is given by the symbol $^{A}_{Z}X$, where X is the element's symbol from the periodic table, Z is the atomic number (the number of protons in the nucleus), and A is the mass number (the number of protons and neutrons in the nucleus). Often, Z is omitted because the element symbol gives us the same information about the identity of the element. To determine the number of neutrons in the nucleus for each named isotope, we subtract Z (number of protons) from A (mass number). If the elements are neutral (no charge), the number of electrons equals the number of protons in the nucleus.

Solve

	Atom	Mass Number	Atomic Number = Number of Protons	Number of Neutrons = Mass Number – Atomic Number	Number of Electrons = Number of Protons
(a)	^{11}B	11	5	6	5
(b)	^{19}F	19	9	10	9
(c)	^{131}I	131	53	78	53
(d)	^{222}Rn	222	86	136	86

Think About It

As the atom gets more massive, the ratio of neutrons to protons gets larger. For fluorine the ratio is 10/9 = 1.1, but for radon the ratio is 136/86 = 1.6. We will see later that as the nucleus adds more protons, it needs more neutrons as "glue" to keep the nucleus stable.

2.27. **Collect and Organize**

After calculating the neutron-to-proton ratio for ^{4}He, ^{23}Na, ^{59}Co, and ^{197}Au, we are to comment on how the ratio changes as Z increases.

Analyze

To calculate the neutron-to-proton ratios, we need to determine, from the mass number and the atomic number, the number of protons and neutrons for each element:

^{4}He, atomic number 2, has 2 protons and 2 neutrons
^{23}Na, atomic number 11, has 11 protons and 12 neutrons
^{59}Co, atomic number 27, has 27 protons and 32 neutrons
^{197}Au, atomic number 79, has 79 protons and 118 neutrons

Solve

The neutron-to-proton ratio for each element is

(a) ^{4}He, 2/2 = 1.00
(b) ^{23}Na, 12/11 = 1.09
(c) ^{59}Co, 32/27 = 1.19
(d) ^{197}Au, 118/79 = 1.49

As the atomic number (Z) increases, the neutron-to-proton ratio increases.

Think About It

More neutrons are required to stabilize nuclei with more protons because of the strong repulsive forces between the positively charged protons. Neutrons bring added strong nuclear force to the nucleus to stabilize it.

2.28. **Collect and Organize**

After calculating the neutron-to-proton ratios for group 15 nuclei ^{14}N, ^{31}P, ^{75}As, ^{121}Sb, and ^{123}Sb, we are to comment on how the ratio changes as Z increases.

Analyze

To calculate the neutron-to-proton ratios, we need to determine, from the mass number and the atomic number, the number of protons and neutrons for each element:

^{14}N, atomic number 7, has 7 protons and 7 neutrons
^{31}P, atomic number 15, has 15 protons and 16 neutrons
^{75}As, atomic number 33, has 33 protons and 42 neutrons
^{121}Sb, atomic number 51, has 51 protons and 70 neutrons
^{123}Sb, atomic number 51, has 51 protons and 72 neutrons

Solve

The neutron-to-proton ratio for each element is

(a) ^{14}N, 7/7 = 1.00
(b) ^{31}P, 16/15 = 1.07

(c) ^{75}As, 42/33 = 1.27
(d) ^{121}Sb, 70/51 = 1.37
(e) ^{123}Sb, 72/51 = 1.41
As the atomic number (Z) increases, the neutron-to-proton ratio increases.

Think About It
Both isotopes of antimony, ^{121}Sb and ^{123}Sb, are present in nature in comparable abundances (57% and 43%, respectively).

2.29. **Collect and Organize**
To fill in the table, we have to consider how the numbers of nuclear particles relate to one another. We also need to recall how the symbols for the isotopes are written. From the table, it is apparent that we have to work backward in some cases from the number of electrons or protons and mass number to get the element symbol.

Analyze
An isotope is given by the symbol $^{A}_{Z}\mathrm{X}$, where X is the element's symbol from the periodic table, Z is the atomic number (the number of protons in the nucleus), and A is the mass number (the number of protons and neutrons in the nucleus). We can determine the number of neutrons in the nucleus for the isotopes by subtracting Z (number of protons) from A (mass number). If the elements are neutral (no charge), the number of electrons equals the number of protons in the nucleus. Except for the first element, we will assume that the rest of the elements are neutral; otherwise, there would be different possibilities for the numbers of protons and electrons.

Solve

Symbol	^{23}Na	^{89}Y	^{118}Sn	^{197}Au
Number of Protons	11	39	50	79
Number of Neutrons	12	50	68	118
Number of Electrons	11	39	50	79
Mass Number	23	89	118	197

Think About It
Because the nuclear particles are all related to one another, either we can work from the isotope symbol to find the number of protons, neutrons, and electrons for a particular isotope, or we can work from the mass number and the number of electrons or protons to determine the number of neutrons and write the element symbol.

2.30. **Collect and Organize**
To fill in the table, we have to consider how the numbers of nuclear particles relate to one another. We also need to recall how the symbols for the isotopes are written. From the table, it is apparent that we have to work backward in some cases from the number of electrons or protons and the mass number to get the element symbol.

Analyze
An isotope is given by the symbol $^{A}_{Z}\mathrm{X}$, where X is the element's symbol from the periodic table, Z is the atomic number (the number of protons in the nucleus), and A is the mass number (the number of protons and neutrons in the nucleus). We can determine the number of neutrons in the nucleus for the isotopes by subtracting Z (number of protons) from A (mass number). If the elements are neutral (no charge), the number of electrons equals the number of protons in the nucleus. Except for the first element, we will assume that the rest of the elements are neutral, otherwise there would be different possibilities for the numbers of protons and electrons.

Solve

Symbol	^{27}Al	^{98}Mo	^{143}Nd	^{238}U
Number of Protons	13	42	60	92
Number of Neutrons	14	56	83	146
Number of Electrons	13	42	60	92
Mass Number	27	98	143	238

Think About It
Because the nuclear particles are all related to one another, either we can work from the isotope symbol to find the number of protons, neutrons, and electrons for a particular isotope, or we can work from the mass number and the number of electrons or protons to determine the number of neutrons and write the element symbol.

2.31. **Collect and Organize**

An isotope of a monatomic ion can be given the symbol ${}^{A}_{Z}X^{n}$, where X is the element's symbol from the periodic table, *Z* is the atomic number (the number of protons in the nucleus), *A* is the mass number (the number of protons and neutrons in the nucleus), and *n* is the charge on the species.

Analyze
If we are given the number of protons in the nucleus, the element can be identified from the periodic table. We can determine the mass number for the isotope by adding the protons and the neutrons in the nucleus. We can determine the number of neutrons or protons by subtracting *Z* (number of protons) or the number of neutrons, respectively, from *A* (mass number). We can account for the charge on the species by adding electrons (to form a negatively charged ion) or by subtracting electrons (to form a positively charged ion).

Solve

Symbol	${}^{37}Cl^{-}$	${}^{23}Na^{+}$	${}^{81}Br^{-}$	${}^{226}Ra^{2+}$
Number of Protons	17	11	35	88
Number of Neutrons	20	12	46	138
Number of Electrons	18	10	36	86
Mass Number	37	23	81	226

Think About It
To form a singly charged ion, there has to be one more electron (for a negative charge) or one fewer electron (for a positive charge) than the number of protons in the nucleus. For a doubly charged ion, we add or take away two electrons.

2.32. **Collect and Organize**

An isotope of a monatomic ion can be given the symbol ${}^{A}_{Z}X^{n}$, where X is the symbol of the element from the periodic table, *Z* is the atomic number (the number of protons in the nucleus), *A* is the mass number (the number of protons and neutrons in the nucleus), and *n* is the charge on the species.

Analyze
If we are given the number of protons in the nucleus, the element can be identified from the periodic table. We can determine the mass number by adding the numbers of protons and neutrons. We can determine the number of neutrons or protons in the nucleus by subtracting *Z* (number of protons) or the number of neutrons, respectively, from *A* (mass number). We can account for the charge on the species by adding electrons (to form a negatively charged ion) or by subtracting electrons (to form a positively charged ion).

Solve

Symbol	${}^{137}Ba^{2+}$	${}^{64}Zn^{2+}$	${}^{32}S^{2-}$	${}^{90}Zr^{4+}$
Number of Protons	56	30	16	40
Number of Neutrons	81	34	16	50
Number of Electrons	54	28	18	36
Mass Number	137	64	32	90

Think About It
These are all multiply charged ions. For anions, we add electrons; for cations, we remove electrons.

2.33. **Collect and Organize**

Knowing that Mendeleev labeled his groups on the left side of the periodic table on the basis of the formulas of the compounds they formed with oxygen, we are to assign his labels to groups 2, 3, and 4 on the modern periodic table.

Analyze
Groups 2, 3, and 4 have cations with charges of 2+, 3+, and 4+, respectively. Oxygen forms an anion of 2– charge. The oxygen compounds that would form would balance the positive cation charge with the negative anion charge by combining the elements in that group with oxygen in whole-number ratios.

Solve
Group 2, with a 2+ charge, would form RO with oxygen; group 3, with a 3+ charge, would form an R_2O_3 compound with oxygen; and group 4, with a 4+ charge, would form an RO_2 compound with oxygen.

Think About It
Mendeleev's brilliant insight was to classify the elements based on similar chemical behaviors.

2.34. **Collect and Organize**
Knowing that Mendeleev labeled his groups on the right side of the periodic table on the basis of the formulas of the compounds they formed with hydrogen, we are to assign modern periodic table labels to his groups of HR, H_2R, and H_3R.

Analyze
Because hydrogen has a charge of 1+, for those compounds we are looking for groups that have typical charges of their anions of 1–, 2–, and 3–.

Solve
Group 17, with a 1– charge, would form HR compounds with hydrogen; group 16, with a 2– charge, would form H_2R compounds with hydrogen; and group 15, with a 3– charge, would form H_3R compounds with hydrogen.

Think About It
Representative compounds with hydrogen from those groups are hydrochloric acid (HCl), water (H_2O), and ammonia (NH_3).

2.35. **Collect and Organize**
We are asked why Mendeleev did not leave spaces for the noble gases in his periodic system.

Analyze
The noble gases are characterized by their remarkable unreactivity. Unreactive elements can be quite unnoticeable because they do not form compounds with other elements. The noble gases were discovered starting in 1894. Mendeleev proposed his periodic system in 1872.

Solve
The noble gases were not discovered until after Mendeleev put together his periodic table. He could not have predicted the existence of the noble gases at the time since (a) none of them were isolated and characterized on the basis of their reactivity (or lack thereof) to indicate their presence in nature and (b) he arranged the elements in order of increasing mass, not atomic number. If he had been aware of atomic numbers as characteristic of the elements, he would have noticed that the atomic numbers for the noble gases were missing as a column in his table.

Think About It
The noble gases are monatomic, are colorless and odorless, and have a remarkably narrow liquid range (their boiling points and melting points are close together).

2.36. **Collect and Organize**
In this question we are asked to describe how the periodic table would look if the A- and B-labeled columns were stacked.

Analyze
For this, we will stack elements in the *n*A column above that of the *n*B column.

Solve

H									
Li	Be	B	C	N	O	F	Ne		
Na	Mg	Al	Si	P	S	Cl	Ar		
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni
Cu	Zn	Ga	Ge	As	Se	Br	Kr		
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd
Ag	Cd	In	Sn	Sb	Te	I	Xe		
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt
Au	Hg	Tl	Pb	Bi	Po	At	Rn		
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds
Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og		

This is nearly the same way that Mendeleev arranged his periodic table based on masses and similar properties of the elements.

Think About It

This other form loses a lot of information about what we now know about the periodic nature of the elements. In particular, it loses information about the electronic structure (how the electrons populate orbitals) which you will learn in Chapter 3.

2.37. **Collect and Organize**

Knowing that the explosive TNT contains second-row elements in groups 14, 15, and 16 as well as hydrogen, we are to name those particular elements.

Analyze

The second-row elements start with lithium and end at neon.

Solve

The elements in TNT besides hydrogen are carbon (group 14), nitrogen (group 15), and oxygen (group 16).

Think About It

Many explosives have the same elements. The powerful C4 explosive is composed mainly of the explosive RDX_2, which has the chemical formula $C_3H_6N_6O_6$.

2.38. **Collect and Organize**

Knowing that the chemical weapon phosgene contains a second-row element in group 16 and a third-row element in group 17 as well as carbon, we are to name those particular elements and give their atomic numbers.

Analyze

The second-row elements start with lithium and end at neon. The third-row elements start at sodium and end at argon.

Solve

The elements in phosgene besides carbon are oxygen (group 16), with atomic number 8, and chlorine (group 17), with atomic number 17.

Think About It

Phosgene is a small molecule ($COCl_2$) that is highly toxic and causes damage to the lungs, leading to suffocation when inhaled.

2.39. **Collect and Organize**

Given information about the elements used in catalytic converters (their group numbers and relationship to one another in the periodic table), we are asked to name the elements.

Analyze
In examining the clues in parts a–c, we see that these elements will all be transition metals located in the middle of the periodic table.

Solve
(a) The group 10 transition metal element in the fifth row of the periodic table (Rb–Xe) is palladium (Pd).
(b) The transition metal element with one fewer proton than Pd is in group 9 and is rhodium (Rh).
(c) The transition metal element with 32 more protons than Pd is platinum (Pt) with 78 protons.

Think About It
These metals are generally as expensive or more so than gold. In May, 2020 gold is priced at about \$1700/oz, with platinum priced less (about \$820/oz), but rhodium and palladium more expensive (\$6500/oz and \$1900/oz, respectively).

2.40. **Collect and Organize**
We are asked to name the fourth-row halogen that can be used as an alternative to chlorine as a disinfectant.

Analyze
The fourth-row elements start at potassium and end at krypton. The halogens are the elements in group 17.

Solve
The fourth-row halogen element is bromine.

Think About It
Iodine, the fifth-row halogen, also can be used as a disinfectant or antiseptic.

2.41. **Collect and Organize**
We are to count the metallic elements in the third row of the periodic table.

Analyze
The third row in the periodic table begins at sodium and ends at argon. The metalloids begin at silicon, and the nonmetals begin with phosphorus.

Solve
Three metallic elements are in the third row in the periodic table: sodium (Na), magnesium (Mg), and aluminum (Al).

Think About It
It's easy to follow the color scheme in the periodic table of the elements on the inside front cover of the text: Metals are tan, nonmetals are blue, and metalloids are green.

2.42. **Collect and Organize**
From the elements in the third row of the periodic table, we are to identify the element that has the chemical properties of a nonmetal but is more like a metal in its physical properties.

Analyze
The third row in the periodic table begins at sodium and ends at argon. The given characteristics mean that this particular element is between a metal and a nonmetal and is therefore one of the semimetals.

Solve
Silicon (Si) is the element described.

Think About It
As a metalloid, silicon has some, but not all, the properties of a metal but has a structure more like that of a nonmetal. Silicon is an important material in semiconducting computer chips.

2.43. **Collect and Organize**

We define *weighted average* for this question.

Analyze

An average is a number that expresses the middle of the data (here, for various masses of atoms or isotopes).

Solve

A weighted average takes into account the proportion of each value in the group of values to be averaged. For example, the average of 2, 2, 2, and 5 would be computed as (2 + 2 + 2 + 5)/4 = 2.75. That average shows the heavier weighting toward the values of 2.

Think About It

Because isotopes for any element are not equally present but have a range of natural abundances, the masses in the periodic table of the elements are calculated as weighted averages.

2.44. **Collect and Organize**

Given that the abundance of the two isotopes of an element are both 50%, we are to express the average atomic mass of the element if the mass of isotope X is m_X and the mass of isotope Y is m_Y.

Analyze

Because each isotope is present in exactly 50% abundance, the average atomic mass will be the simple average of the two masses of isotopes X and Y.

Solve

$$\text{average atomic mass} = \frac{m_X + m_Y}{2}$$

Think About It

Some elements in nature have just one naturally occurring isotope. Can you find some examples in Appendix 3.3?

2.45. **Collect and Organize**

For calculating the masses of binary ionic compounds (such as NaCl), we are to explain why we use the average masses of the neutral atoms rather than the masses of the ions.

Analyze

For a cation, the actual mass would be the mass of the neutral atom minus the mass of the one or more electrons lost; for an anion, the actual mass would be the atomic mass plus the mass of the electrons gained.

Solve

First, the mass of the electron is small in comparison with the mass of a neutral atom, so the mass of an ion is exceedingly close to the mass of the neutral atom. But, more importantly, in a neutral salt, the number of electrons gained to form the anion and the number of electrons lost to form the cation are equal, so no electrons have been lost or gained in total compared with the number of electrons in the neutral atoms.

Think About It

In NaCl, sodium (with 11 electrons) loses one electron to be Na^+ (now with 10 electrons), and chlorine (with 17 electrons) gains one electron to be Cl^- (with 18 electrons). The total count of electrons stays at 28. Can you do the electron count for $MgCl_2$?

2.46. **Collect and Organize**

We have to consider the concept of weighted average atomic mass to answer this question, and we'll need to look up atomic masses for B, Li, and N from inside the front cover of the textbook.

Analyze

We are asked to compare two isotopes and their weighted average mass. If the lighter isotope is more abundant, the average atomic mass will be less than the average if both isotopes were equally abundant. If the heavier isotope is more abundant, the average atomic mass will be greater than the simple average of the two isotopes. We are given the mass number for the isotopes as part of the isotope symbol, and we will take that as the mass of that isotope in atomic mass units.

Solve

(a) The simple average atomic mass for ^{10}B and ^{11}B would be 10.5 u. The actual average mass (10.811 u) is greater than this; therefore, ^{11}B is more abundant.

(b) The simple average atomic mass for ^{6}Li and ^{7}Li would be 6.5 u. The actual average mass (6.941 u) is greater than this; therefore, ^{7}Li is more abundant.

(c) The simple average atomic mass for ^{14}N and ^{15}N would be 14.5 u. The actual average mass (14.007 u) is less than this; therefore, ^{14}N is more abundant.

Think About It

This is a quick question to answer for elements such as boron, lithium, and nitrogen that have the dominance of only two isotopes in terms of their abundance. Answering the same question for elements with more than two stable isotopes in relatively high abundances is a little harder.

2.47. **Collect and Organize**

We are to determine which isotope of argon is most abundant, ^{36}Ar, ^{38}Ar, or ^{40}Ar.

Analyze

To help here, it is useful to know from the periodic table that the average atomic mass of argon is 39.948 u.

Solve

Only if the highest mass isotope were most abundant would the average mass of argon be 39.948 u. Therefore, ^{40}Ar is the most abundant isotope of argon.

Think About It

Table A3.3 shows these natural abundances: ^{40}Ar, 99.600%; ^{38}Ar, 0.063%; and ^{36}Ar, 0.337%.

2.48. **Collect and Organize**

Given that the two most common isotopes of hydrogen have 0 or 1 neutron (i.e., are ^{1}H and ^{2}H) and those of oxygen have 8 or 10 neutrons (i.e., are ^{16}O and ^{18}O), we are to determine the possible values for the molecular weight (in u) of a single water molecule. We'll need the masses for subatomic particles (Table 2.1).

Analyze

The possible water molecules are: $^{1}H_2{}^{16}O$, $^{1}H^{2}H^{16}O$, $^{2}H_2{}^{16}O$, $^{2}H_2{}^{18}O$, $^{1}H^{2}H^{18}O$, and $^{1}H_2{}^{18}O$. All of these have the same number of protons, 10, for which we can find the mass by multiplying by the mass of the proton (1.00728 u). Each of these neutral molecules has 10 electrons, for which we can find the mass by multiplying by the mass of the electron (5.48580×10^{-4} u). Finally, for each molecule we will calculate the total mass of neutrons using the mass of a neutron (1.00866 u). We will then sum these (mass of protons, electrons, and neutrons for each molecule) to get the atomic weight of each type of water molecule.

Solve

$^{1}H_2{}^{16}O$ has 10 protons, 10 electrons, and 8 neutrons:

$$\left(10\text{ protons}\times\frac{1.00728\text{ u}}{\text{proton}}\right)+\left(10\text{ electrons}\times\frac{5.48580\times10^{-4}\text{ u}}{\text{electron}}\right)+\left(8\text{ neutrons}\times\frac{1.00866\text{ u}}{\text{proton}}\right)=18.1476\text{ u}$$

$^{1}H^{2}H^{16}O$:

$$\left(10\text{ protons}\times\frac{1.00728\text{ u}}{\text{proton}}\right)+\left(10\text{ electrons}\times\frac{5.48580\times10^{-4}\text{ u}}{\text{electron}}\right)+\left(9\text{ neutrons}\times\frac{1.00866\text{ u}}{\text{proton}}\right)=19.1562\text{ u}$$

$^{2}H_2{}^{16}O$:

$$\left(10\ \text{protons}\times\frac{1.00728\ \text{u}}{\text{proton}}\right)+\left(10\ \text{electrons}\times\frac{5.48580\times10^{-4}\ \text{u}}{\text{electron}}\right)+\left(10\ \text{neutrons}\times\frac{1.00866\ \text{u}}{\text{proton}}\right)=20.1649\ \text{u}$$

$^{2}H_2{}^{18}O$:

$$\left(10\ \text{protons}\times\frac{1.00728\ \text{u}}{\text{proton}}\right)+\left(10\ \text{electrons}\times\frac{5.48580\times10^{-4}\ \text{u}}{\text{electron}}\right)+\left(12\ \text{neutrons}\times\frac{1.00866\ \text{u}}{\text{proton}}\right)=22.1822\ \text{u}$$

$^{1}H^{2}H^{18}O$:

$$\left(10\ \text{protons}\times\frac{1.00728\ \text{u}}{\text{proton}}\right)+\left(10\ \text{electrons}\times\frac{5.48580\times10^{-4}\ \text{u}}{\text{electron}}\right)+\left(11\ \text{neutrons}\times\frac{1.00866\ \text{u}}{\text{proton}}\right)=21.1735\ \text{u}$$

$^{1}H_2{}^{18}O$:

$$\left(10\ \text{protons}\times\frac{1.00728\ \text{u}}{\text{proton}}\right)+\left(10\ \text{electrons}\times\frac{5.48580\times10^{-4}\ \text{u}}{\text{electron}}\right)+\left(10\ \text{neutrons}\times\frac{1.00866\ \text{u}}{\text{proton}}\right)=20.1649\ \text{u}$$

Think About It

Hydrogen has three common isotopes and because each has a very different mass, they have special names: ^{1}H is protium, ^{2}H is deuterium, and ^{3}H is tritium.

2.49. **Collect and Organize**

In this question we are asked to explain the observation that the average atomic mass of platinum is 195.08 u, whereas the natural abundance of ^{195}Pt is only 33.8%.

Analyze

The average atomic mass on the periodic table is the weighted average of all the masses of the naturally occurring isotopes for that element. Only if an element has one naturally occurring element will the average atomic mass match the mass of that isotope.

Solve

We are given that the ^{195}Pt isotope is not 100% abundant. Therefore, we must conclude that the other isotopes with masses greater than 195 u have natural abundances in equal proportion to those isotopes with masses lower than 195 so that the weighted average atomic mass calculates to 195.08 u.

Think About It

For example, platinum might have three isotopes, each in 33% abundance: ^{194}Pt, ^{195}Pt, and ^{196}Pt.

2.50. **Collect and Organize**

In this question we are asked to determine whether we can conclude that europium has only one stable isotope on the basis of the information that its atomic mass is only 0.04 u off from a whole number.

Analyze

The average atomic mass on the periodic table is the weighted average of all the masses of the naturally occurring isotopes for that element. Only if an element has one naturally occurring element will the average atomic mass match the mass of that isotope. The mass of an isotope is the sum of the masses of all its protons, neutrons, and electrons.

Solve

No, we cannot conclude from this information that only one stable isotope of europium exists. Heavier and lighter isotopes might "cancel" each other in the calculation of the average atomic mass, or other stable isotopes might be in very small abundances and so they do not contribute significantly to the weighted atomic mass.

Think About It
In fact, europium has two naturally occurring isotopes: ^{151}Eu, which is 52% abundant, and ^{153}Eu, which is about 48% abundant.

2.51. **Collect and Organize**

For a sample of hydrogen gas (H_2) in which ^{1}H is 99.99745% abundant and ^{2}H is 0.00255% abundant, we are to calculate the average atomic mass of hydrogen and compare it to the mass of H in the periodic table in the textbook.

Analyze
To calculate the average atomic mass, we have to consider the relative abundances according to the following formula:

$$m_X = a_1m_1 + a_2m_2 + a_3m_3 + \cdots$$

where a_n refers to the abundance of isotope n and m_n refers to the mass of isotope n. If the relative abundances are given as percentages, the value we use for a_n in the formula is the percentage divided by 100. The masses of ^{1}H and ^{2}H are given in the statement of the problem.

Solve

$$m_H = (0.9999745 \times 1.007825 \text{ u}) + (0.0000255 \times 2.014102 \text{ u}) = 1.007851 \text{ u}$$

This value matches that given in the periodic table (1.0079 u) to five significant figures.

Think About It
Because the ^{1}H isotope is in much, much greater abundance than ^{2}H, we expect the average atomic mass of hydrogen in any sample of H_2 to be very close to 1 u rather than 2 u.

2.52. **Collect and Organize**

For a boron-containing sample of water in which ^{11}B is 81.07% abundant and ^{10}B is 18.93% abundant, we are to calculate the average atomic mass of boron and compare it to the mass of B in the periodic table in the textbook.

Analyze
To calculate the average atomic mass, we have to consider the relative abundances according to the following formula:

$$m_X = a_1m_1 + a_2m_2 + a_3m_3 + \cdots$$

where a_n refers to the abundance of isotope n and m_n refers to the mass of isotope n. If the relative abundances are given as percentages, the value we use for a_n in the formula is the percentage divided by 100. The masses of ^{11}B and ^{10}B are given in the statement of the problem.

Solve

$$m_B = (0.8107 \times 11.0093 \text{ u}) + (0.1893 \times 10.0129 \text{ u}) = 10.82 \text{ u}$$

This value nearly matches that on the periodic table (10.811 u) to four significant figures.

Think About It
Because the ^{11}B isotope is in much greater abundance than the ^{10}B isotope, we expect the average atomic mass to be closer to 11 u rather than 10 u.

2.53. **Collect and Organize**

Here we are asked to find out whether the mass of magnesium on Mars is the same as here on Earth. We are given the masses of each of the three isotopes of Mg in the Martian sample. Once we calculate the weighted average for Mg for the Martian sample, we can compare it with the average mass for Mg found on Earth.

Analyze
To calculate the average atomic mass, we have to consider the relative abundances according to the following formula:

$$m_X = a_1m_1 + a_2m_2 + a_3m_3 + \cdots$$

where a_n refers to the abundance of isotope n and m_n refers to the mass of isotope n. If the relative abundances are given as percentages, the value we use for a_n in the formula is the percentage divided by 100.

Solve

For the average atomic mass of magnesium in the Martian sample

$$\begin{aligned} m_{\text{Mg}} &= (0.7870 \times 23.9850\ \text{u}) + (0.1013 \times 24.9858\ \text{u}) + (0.1117 \times 25.9826\ \text{u}) \\ &= 24.31\ \text{u} \end{aligned}$$

The average mass of Mg on Mars is the same as here on Earth.

Think About It

The mass of Mg on Mars should be close to the same value as on Earth; the magnesium on both planets arrived in the solar system via the same ancient stardust.

2.54. **Collect and Organize**

Given the exact mass for each isotope of lithium and its abundance, we are to calculate the average mass for lithium and compare that value with the mass found in the periodic table.

Analyze

The weighted atomic mass is computed according to the formula

$$m_{\text{X}} = a_1m_1 + a_2m_2 + a_3m_3 + \cdots$$

where a_n refers to the abundance of isotope n and m_n refers to the mass of isotope n. If the relative abundances are given as percentages, the value we use for a_n in the formula is the percentage divided by 100.

Solve

$$m_{\text{Li}} = (0.0773 \times 6.015123\ \text{u}) + (0.9227 \times 7.016003\ \text{u}) = 6.939\ \text{u}$$

The atomic mass of Li in the periodic table is 6.941 u. These values are close.

Think About It

We expect the average mass computed here to closely match the value in the periodic table because the periodic table value also is the weighted average for the atomic mass.

2.55. **Collect and Organize**

In this problem, we again use the concept of weighted average atomic mass, but here we have to work backward from the average mass to find the exact mass of the ^{48}Ti isotope, one of the five stable Ti isotopes.

Analyze

We can use the formula for finding the weighted average atomic mass, but this time our unknown quantity is one of the isotope masses. Here,

$$m_{\text{Ti}} = a_{^{46}\text{Ti}}m_{^{46}\text{Ti}} + a_{^{47}\text{Ti}}m_{^{47}\text{Ti}} + a_{^{48}\text{Ti}}m_{^{48}\text{Ti}} + a_{^{49}\text{Ti}}m_{^{49}\text{Ti}} + a_{^{50}\text{Ti}}m_{^{50}\text{Ti}}$$

Solve

$$\begin{aligned} 47.867\ \text{u} &= (0.0825 \times 45.9526\ \text{u}) + (0.0744 \times 46.9518\ \text{u}) + (0.7372 \times m_{^{48}\text{Ti}}) \\ &\quad + (0.0541 \times 48.94787\ \text{u}) + (0.0518 \times 49.94479\ \text{u}) \\ m_{^{48}\text{Ti}} &= 47.95\ \text{u} \end{aligned}$$

Think About It

That answer makes sense because the exact mass of ^{48}Ti should be close to 48 u.

2.56. **Collect and Organize**

In this problem, we again use the concept of weighted average atomic mass, but here we are asked to work backward from the average mass to find the exact mass of the ^{92}Zr isotope.

Analyze
We can use the formula for finding the weighted average atomic mass, only this time our unknown quantity is one of the isotope masses. Here,

$$m_{Zr} = a_{^{90}Zr}m_{^{90}Zr} + a_{^{91}Zr}m_{^{91}Zr} + a_{^{92}Zr}m_{^{92}Zr} + a_{^{94}Zr}m_{^{94}Zr} + a_{^{96}Zr}m_{^{96}Zr}$$

Solve
$$91.224\text{ u} = (0.5145 \times 89.905\text{ u}) + (0.1122 \times 90.906\text{ u}) + (0.1715 \times m_{^{92}Zr}) + (0.1738 \times 93.906\text{ u}) + (0.0280 \times 95.908)$$

$$m_{^{92}Zr} = 91.906\text{ u}$$

Think About It
That answer makes sense because the exact mass of ^{92}Ti should be close to 92 u.

2.57. **Collect and Organize**

From the formulas for four ionic compounds, NaI, CaS, Al_2O_3, and NH_4Cl, we are to calculate the masses of the formula units.

Analyze
For each formula we will sum the masses of the elements, making sure that we also account for the number of a particular element present in the formula.

Solve
(a) NaI: 22.990 u + 126.90 u = 149.89 u
(b) CaS: 40.078 u + 32.065 u = 72.143 u
(c) Al_2O_3: 2(26.982 u) + 3(15.999 u) = 101.961 u
(d) NH_4Cl: 14.007 u + 4(1.0079 u) + 35.453 u = 53.492 u

Think About It
In determining the formula mass, use as many significant figures in your calculation as listed in the periodic table. Resist the temptation to round up or down, which would make the calculation for the mass less accurate.

2.58. **Collect and Organize**

From the formulas for four ionic compounds, LiCl, TiO_2, $MgCO_3$, and $CaHPO_4$, we are to calculate the masses of the formula units.

Analyze
For each formula we will sum the masses of the elements, making sure that we also account for the number of a particular element present in the formula.

Solve
(a) LiCl: 6.941 u + 35.453 u = 42.394 u
(b) TiO_2: 47.867 u + 2(15.999 u) = 79.865 u
(c) $MgCO_3$: 24.305 u + 12.011 u + 3(15.999 u) = 84.313 u
(d) $CaHPO_4$: 40.078 u + 1.0079 u + 30.974 u + 4(15.999 u) = 136.056 u

Think About It
In determining the formula mass, use as many significant figures in your calculation as listed in the periodic table. Resist the temptation to round up or down, which would make the calculation for the mass less accurate.

2.59. **Collect and Organize**

For each given formula, we are to determine the number of carbon atoms in each molecule.

Analyze
Each C in the formula stands for a C atom, and any subscript after the C atom is the number of carbons in that part of the molecule.

Solve
(a) CO_2 has 1 C atom per molecule
(b) $HOCH_2CH_2OH$ has 2 atoms per molecule
(c) CH_3COOH has 2 atoms per molecule
(d) C_2H_5OH has 2 atoms per molecule

Think About It
Those molecules are all organic molecules, and writing their fully condensed formulas as $C_aH_bN_cO_d$ followed by other elements, if present, is customary. Formulas like $HOCH_2CH_2OH$ or C_2H_5OH show how atoms are connected in the molecule.

2.60. **Collect and Organize**
For each given formula in Problem 2.59, we are to determine the number of oxygen atoms in each.

Analyze
The subscript in each formula after the O atom is the number of oxygens in the molecule.

Solve
(a) CO_2, has 2 atoms per molecule
(b) $HOCH_2CH_2OH$, has 2 atoms per molecule
(c) CH_3COOH, has 2 atoms per molecule
(d) C_2H_5OH, has 1 atom per molecule

Think About It
Those molecules are all organic molecules, and writing their formula as $C_aH_bN_cO_d$ followed by other elements, if present, is customary.

2.61. **Collect and Organize**
For a list of five compounds, we are to rank them in order of increasing mass. First, we have to determine their molecular masses.

Analyze
When we use the masses on the periodic table to calculate the molecular masses, we obtain the following:
(a) CS_2 = 76.14 u
(b) HI = 127.91 u
(c) NO_2 = 46.01 u
(d) HClO = 52.46 u
(e) C_4H_{10} = 58.12 u

Solve
In order of increasing molecular mass: (c) NO_2 < (d) HClO < (e) C_4H_{10} < (a) CS_2 < (b) HI.

Think About It
In this problem fewer significant figures were necessary for the molecular masses because we were going to compare masses, and the masses were not likely to be too close together to warrant more precise values.

2.62. **Collect and Organize**
For a list of five compounds, we are to rank them in order of decreasing mass. First, we need to determine their molecular masses.

Analyze
When we use the masses in the periodic table to calculate the molecular masses, we obtain the following:
(a) H_2O = 18.01 u
(b) CO_2 = 44.009 u
(c) N_2O = 44.013 u
(d) CH_2Cl_2 = 84.93 u
(e) BF_3 = 67.81 u

Solve
In order of decreasing molecular mass: (d) CH_2Cl_2 > (e) BF_3 > (c) N_2O > (b) CO_2 > (a) H_2O.

Think About It
In this problem we needed five significant figures for the mass to discriminate the order between CO_2 and N_2O.

2.63. **Collect and Organize**
For describing a collection of atoms or molecules, we are asked why using the unit *dozen* to express the number of atoms or molecules we have might not be a good idea.

Analyze
A dozen is 12 objects and therefore a relatively small group.

Solve
Although a dozen is a convenient and recognizable unit for donuts and eggs, it is too small a unit to express the very large number of atoms, ions, or molecules present in a mole.

$$\frac{6.022\times10^{23}\ \text{atoms}}{\text{mole}}\times\frac{1\ \text{dozen}}{12\ \text{atoms}}=5.02\times10^{22}\ \text{dozen/mole}$$

Think About It
The mole (6.022×10^{23}) is a much more convenient unit to express the number of atoms or molecules in a sample.

2.64. **Collect and Organize**
We are asked whether equal masses of two isotopes of an element contain the same number of atoms.

Analyze
The number of atoms present for a given mass is dependent on the molar mass of the substance. Isotopes with higher molar masses will have fewer atoms for a given mass of the element than those with a lower isotopic mass.

Solve
No, given samples of equal masses, the isotope with the higher molar mass will contain fewer atoms than the isotope of lower molar mass.

Think About It
That difference may not be significant, however, since many isotopic molar masses are close to each other.

2.65. **Collect and Organize**
In this exercise, we convert the given number of atoms or molecules of each gas to moles.

Analyze
To convert the number of atoms or molecules to moles, we divide by the Avogadro constant.

Solve
(a) $$\frac{4.4\times10^{14}\ \text{atoms of Ne}}{6.022\times10^{23}\ \text{atoms/mol}}=7.3\times10^{-10}\ \text{mol Ne}$$

(c) $$\frac{2.5\times10^{12}\ \text{molecules of O}_3}{6.022\times10^{23}\ \text{molecules/mol}}=4.2\times10^{-12}\ \text{mol O}_3$$

(b) $\dfrac{4.2\times10^{13}\ \text{molecules of CH}_4}{6.022\times10^{23}\ \text{molecules/mol}} = 7.0\times10^{-11}\ \text{mol CH}_4$ (d) $\dfrac{4.9\times10^{9}\ \text{molecules of NO}_2}{6.022\times10^{23}\ \text{molecules/mol}} = 8.1\times10^{-15}\ \text{mol NO}_2$

Think About It
The trace gas with the most atoms or molecules present also has the most moles present. In that sample of air, the amount of the trace gases decreases in the order Ne > CH_4 > O_3 > NO_2.

2.66. **Collect and Organize**
In this exercise, we convert the number of molecules of each gas found in the sample to moles.

Analyze
To convert the number of molecules to moles, we divide by the Avogadro constant.

Solve
(a) $\dfrac{1.4\times10^{13}\ \text{molecules of H}_2}{6.022\times10^{23}\ \text{molecules/mol}} = 2.3\times10^{-11}\ \text{mol H}_2$ (c) $\dfrac{7.7\times10^{12}\ \text{molecules of N}_2\text{O}}{6.022\times10^{23}\ \text{molecules/mol}} = 1.3\times10^{-11}\ \text{mol N}_2\text{O}$

(b) $\dfrac{1.5\times10^{14}\ \text{atoms of He}}{6.022\times10^{23}\ \text{atoms/mol}} = 2.5\times10^{-10}\ \text{mol He}$ (d) $\dfrac{3.0\times10^{12}\ \text{molecules of CO}}{6.022\times10^{23}\ \text{molecules/mol}} = 5.0\times10^{-12}\ \text{mol CO}$

Think About It
The trace gas with the most atoms or molecules present also has the most moles present. In that sample of air, the amount of the trace gases decreases in the order He > H_2 > N_2O > CO.

2.67. **Collect and Organize**
From the chemical formulas for various nitrogen compounds with oxygen and hydrogen, we are asked to determine how many moles of nitrogen are in 1 mol of each substance.

Analyze
The chemical formula reflects the molar ratios of the elements in the compound. If one atom of nitrogen is in the compound's chemical formula, then 1 mol of nitrogen is in 1 mol of the compound. Likewise, if three atoms of nitrogen are in the chemical formula, 3 mol of nitrogen is present in 1 mol of the substance.

Solve
(a) One atom of nitrogen is in NO; therefore, 1 mol of NO contains 1 mol of nitrogen.
(b) One atom of nitrogen is in NO_2; therefore, 1 mol of NO_2 contains 1 mol of nitrogen.
(c) Two atoms of nitrogen are in NH_4NO_3; therefore, 1 mol of NH_4NO contains 2 mol of nitrogen.
(d) Two atoms of nitrogen are in N_2O_5; therefore, 1 mol of N_2O_5 contains 2 mol of nitrogen.

Think About It
If the question had asked how many moles of oxygen were present in 1 mol of N_2O_5, the answer would be 5 mol of oxygen.

2.68. **Collect and Organize**
From the chemical formulas for various potassium salts, we are asked to determine how many moles of K^+ ions are in 1 mol of each substance.

Analyze
The chemical formula reflects the molar ratios of the elements in the compound. If one atom of potassium is in the compound's chemical formula, then 1 mol of K^+ is in 1 mol of the compound. Likewise, if two atoms of potassium are in the chemical formula, 2 mol of K^+ is present in 1 mol of the substance.

Solve
(a) One atom of potassium is in KBr; therefore, 1 mol of KBr contains 1 mol of K^+.
(b) Two atoms of potassium are in K_2SO_4; therefore, 1 mol of K_2SO_4 contains 2 mol of K^+.

(c) One atom of potassium is in KH_2PO_4; therefore, 1 mol of KH_2PO contains 1 mol of K^+.
(d) Two atoms of potassium are in K_2O; therefore, 1 mol of K_2O contains 2 mol of K^+.

Think About It
The number of potassium ions appearing in those chemical formulas balances the charge on the anions in the salts. For example, two potassium cations are needed to produce the neutral salt of SO_4^{2-}.

2.69. **Collect and Organize**
Given the formulas and the moles of each substance in a pair, we are asked to decide which compound contains more moles of oxygen.

Analyze
To answer this question, we have to take into account the moles of oxygen present in the substance formulas as well as the initial number of moles specified for each substance.

Solve
(a) One mole of MnO_2 contains 2 mol of oxygen, and 1 mol of Fe_2O_3 contains 3 mol of oxygen. Therefore, the one mole of Fe_2O_3 contains more moles of oxygen.
(b) One and one-half moles of CO_2 contains 3 mol of oxygen, and 1 mol of N_2O_5 contains 5 mol of oxygen. Therefore, the 1 mol of N_2O_5 contains more moles of oxygen.
(c) Three moles of NO contain 3 mol of oxygen, and 2 mol of SO_2 contains 4 mol of oxygen. Therefore, the 2 mol of SO_2 contains more oxygen.

Think About It
We cannot decide which substance has more moles of oxygen by comparing only the amounts of the substances present. If that were the case, we would have incorrectly concluded that 3 mol of NO contains more moles of oxygen than 2 mol of SO_2.

2.70. **Collect and Organize**
Given the formulas and the moles of each substance in a pair, we are asked to decide which compound contains more moles of oxygen.

Analyze
To answer this question, we have to take into account the moles of oxygen present in the substance formulas as well as the initial number of moles specified for each substance.

Solve
(a) One mole of N_2O contains 1 mol of oxygen, and 1 mol of NO_2 contains 2 mol of oxygen. Therefore, the 1 mol of NO_2 contains more oxygen.
(b) One mole of H_2SO_4 contains 4 mol of oxygen, and 1.5 mol of H_2SO_3 contains 4.5 mol of oxygen. Therefore, the 1.5 mol of H_2SO_3 contain more moles of oxygen.
(c) One mole of $Ca_3(PO_4)_2$ contains 8 mol of oxygen, and 1.5 mol of $Ca(HCO_3)_2$ contains 9 mol of oxygen. Therefore, the 1.5 mol of $Ca(HCO_3)_2$ contains more oxygen.

Think About It
We cannot decide which substance has more moles of oxygen by comparing only the amounts of substances present.

2.71. **Collect and Organize**
This exercise has us compute the molar masses of various molecular compounds of oxygen.

Analyze
We can find the molar mass of each gas by adding the molar mass of each element from the periodic table, taking into account the number of moles of each atom present in 1 mol of the substance.

Solve

(a) SO_2: 32.065 + 2(15.999) = 64.063 g/mol
(b) O_3: 3(15.999) = 47.997 g/mol
(c) CO_2: 12.011 + 2(15.999) = 44.009 g/mol
(d) N_2O_5: 2(14.007) + 5(15.999) = 108.009 g/mol

Think About It

The three gases SO_2, O_3, and CO_2 have three atoms in their chemical formulas, but each molecule has a different molar mass because they have different constituent atoms.

2.72. **Collect and Organize**

This exercise has us compute the molar masses of various minerals.

Analyze

We can find the molar mass of each mineral by adding the molar mass of each element from the periodic table, taking into account the number of moles of each atom present in 1 mol of the mineral.

Solve

(a) Rhodonite, $MnSiO_3$: 54.938 + 28.086 + 3(15.999) = 131.021 g/mol
(b) Scheelite, $CaWO_4$: 40.078 + 183.84 + 4(15.999) = 287.91 g/mol
(c) Ilmenite, $FeTiO_3$: 55.845 + 47.867 + 3(15.999) = 151.709 g/mol
(d) Magnesite, $MgCO_3$: 24.305 + 12.011 + 3(15.999) = 84.313 g/mol

Think About It

Scheelite, $CaWO_4$, has the highest molar mass of those minerals; therefore, if we had 100 g of each mineral, that mass of scheelite would have the smallest number of moles.

2.73. **Collect and Organize**

This exercise has us compute the molar mass of various flavorings.

Analyze

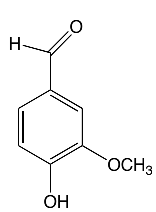
We can find the molar mass of each flavoring by adding the molar mass of each element from the periodic table, taking into account the number of moles of each atom present in 1 mol of the flavoring. Each flavoring contains only carbon (12.011 g/mol), hydrogen (1.0079 g/mol), and oxygen (15.999 g/mol).

Solve

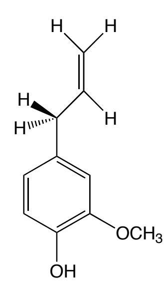
(a) Vanillin, $C_8H_8O_3$: 8(12.011) + 8(1.0079) + 3(15.999) = 152.148 g/mol
(b) Oil of cloves, $C_{10}H_{12}O_2$: 10(12.011) + 12(1.0079) + 2(15.999) = 164.203 g/mol
(c) Anise oil, $C_{10}H_{12}O$: 10(12.011) + 12(1.0079) + 15.999 = 148.204 g/mol
(d) Oil of cinnamon, C_9H_8O: 9(12.011) + 8(1.0079) + 15.999 = 132.161 g/mol

Think About It

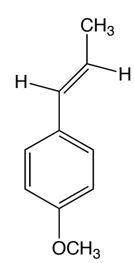
Each flavoring has a distinctive odor and flavor due in part to its different chemical formula. Another factor, however, in differentiating these flavorings is their chemical structure, or the arrangement in which the atoms are attached, as shown by the structures of those flavorings:



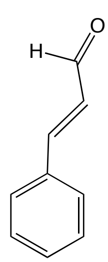
Vanillin



Oil of cloves



Anise oil



Oil of cinnamon

2.74. **Collect and Organize**

This exercise has us compute the molar masses of various sweeteners.

Analyze

We can find the molar mass of each sweetener by adding the molar mass of each element from the periodic table, taking into account the number of moles of each atom present in 1 mol of the sweetener. These flavorings contain carbon (12.011 g/mol), hydrogen (1.0079 g/mol), oxygen (15.999 g/mol), nitrogen (14.007 g/mol), and sulfur (32.065 g/mol).

Solve

(a) Sucrose, $C_{12}H_{22}O_{11}$: 12(12.011) + 22(1.0079) + 11(15.999) = 342.295 g/mol
(b) Saccharin, $C_7H_5O_3NS$: 7(12.011) + 5(1.0079) + 3(15.999) + 14.007 + 32.065 = 183.186 g/mol
(c) Aspartame, $C_{14}H_{18}N_2O_5$: 14(12.011) +18(1.0079) + 2(14.007) + 5(15.999) = 294.305 g/mol
(d) Fructose, $C_6H_{12}O_6$: 6(12.011) + 12(1.0079) + 6(15.999) = 180.155 g/mol

Think About It

Two of those sweeteners are natural (fructose and sucrose), and two are artificial (saccharin and aspartame).

2.75. **Collect and Organize**

We are asked to convert a mass of SiO_2 in quartz in grams to moles.

Analyze

We need the mass of 1 mol of SiO_2 to compute the number of moles in the 7.474 g quartz sample. From the periodic table, we calculate the molar mass to be 60.084 g/mol.

Solve

$$7.474\ \text{g} \times \frac{1\ \text{mol}}{60.084\ \text{g}} = 0.1244\ \text{mol}$$

Think About It

Because the mass of the sample was much lower than the molar mass, the moles of quartz in this sample is quite less than a mole.

2.76. **Collect and Organize**

We are asked to convert a mass of aluminum in kg to moles.

Analyze

We need the mass of 1 mol of aluminum to compute the number of moles of Al in the 2.8 kg sample. From the periodic table, we see that the molar mass of aluminum is 26.982 g/mol. We also must convert the mass of the aluminum in kg to the mass in grams, using the conversion 1 kg = 1000 g.

Solve

$$2.8\ \text{kg} \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{mol}}{26.982\ \text{g}} = 104\ \text{mol or 100 mol (2 significant figures)}$$

Think About It

The number of moles calculated for aluminum is large because we have a large mass of a light element.

2.77. **Collect and Organize**

Given a molar amount of calcium titanate, we are asked to determine the number of moles and mass (in g) of Ca^{2+} ions in the substance.

Analyze

The formula for calcium titanate gives us the number of moles of Ca in the compound. Because the mass of the two missing electrons in the Ca^{2+} cation is negligible, the molar mass of the Ca^{2+} ion is taken to be the same as the molar mass of Ca. We can use that value to determine the mass of Ca^{2+} in the 0.25 mol of calcium titanate.

Solve

Because calcium titanate contains one atom of Ca in its formula, 0.25 mol of $CaTiO_3$ contains 0.25 mol of Ca^{2+} ions. The mass of Ca^{2+} ions in the sample, therefore, is

$$0.25 \text{ mol } Ca^{2+} \times \frac{40.078 \text{ g}}{1 \text{ mol}} = 10 \text{ g } Ca^{2+}$$

Think About It

Our answer makes sense. One-quarter of a mole of Ca^{2+} should give us 40/4, or about 10 g, of Ca in the 0.25 mol of calcium titanate.

2.78. **Collect and Organize**

Given a molar amount of aluminum oxide, we are asked to determine the number of moles and mass of O^{2-} ions in the substance.

Analyze

Aluminum oxide, Al_2O_3, contains 3 mol of oxygen ions per 1 mol as given by its formula. Because the mass of the two additional electrons in the O^{2-} anion is negligible, the molar mass of the O^{2-} anion is taken to be the same as the molar mass of O. We can use that value to determine the mass of O^{2-} in the 0.55 mol of aluminum oxide.

Solve

Because aluminum oxide contains three atoms of O in its formula, 0.55 mol of Al_2O_3 contains $0.55 \times 3 = 1.65$ mol of O^{2-} ions (note that the final answer should be expressed in two significant figures). The mass of O^{2-} ions in the sample, therefore, is

$$1.65 \text{ mol } O^{2-} \times \frac{15.999 \text{ g}}{1 \text{ mol}} = 26 \text{ g}$$

Think About It

Be careful to take into account the number of atoms in the formula of a substance. Here we might have wrongly thought that only 0.55 mol of O^{2-} ions was in the substance and computed a mass that would be one-third the true mass present in 0.55 mol of aluminum oxide.

2.79. **Collect and Organize**

Between two balloons filled with 10.0 g of different gases, we are to choose which balloon has more particles.

Analyze

The balloon with more particles has more moles. The greater number of moles contained in 10.0 g of a gas is for the gas with the lowest molar mass. A gas with a lower molar mass contains more moles in a 10.0 g mass and, therefore, has more moles than a 10.0 g mass of a higher molar mass gas.

Solve

(a) The molar mass of CO_2 is about 44 g/mol, and the molar mass of NO is about 30 g/mol. Therefore, the balloon containing NO has more particles.
(b) The molar mass of CO_2 is 44 g/mol, and the molar mass of SO_2 is 64 g/mol. Therefore, the balloon containing CO_2 has more particles.
(c) The molar mass of O_2 is 32 g/mol, and the molar mass of Ar is 40 g/mol. Therefore, the balloon containing O_2 has more particles.

Think About It

Although we could numerically determine the number of moles of gas in each balloon to make the comparisons in this problem, doing so is unnecessary because we know the relationship between moles and molar mass.

2.80. **Collect and Organize**

If we have equal masses of two salts with different formula weights, which contains more ions?

Analyze

To determine the salt with more ions, we have to take into account not only the formula weight difference between the salts but also the number of ions in the chemical formula for the salt. If both salts contain only one cation and one anion, then the salt sample with the lowest formula mass contains more ions. If the salts contain different numbers of ions, then we have to compute the moles of ions in each case, by assuming a certain amount (in grams) of the salts.

Solve

(a) The formula mass of NaBr is 102.9 g/mol, and that of KCl is 74.6 g/mol. Because each salt has two ions in its formula, the one with the lowest formula mass has more ions for a given mass of salt, KCl.

(b) The molar mass of NaCl is 58.44 g/mol, and the molar mass of $MgCl_2$ is 95.21 g/mol. Because NaCl has two ions in its formula and $MgCl_2$ has three, which has more ions is not immediately obvious. To determine that, assume a convenient number of grams (10 g), convert the grams to moles of ions for each salt, and compare:

$$10 \text{ g NaCl} \times \frac{1 \text{ mol}}{58.44 \text{ g}} \times \frac{2 \text{ ions}}{\text{NaCl formula}} = 0.34 \text{ mol ions in NaCl}$$

$$10 \text{ g MgCl}_2 \times \frac{1 \text{ mol}}{95.21 \text{ g}} \times \frac{3 \text{ ions}}{\text{MgCl}_2 \text{ formula}} = 0.32 \text{ mol ions in MgCl}_2$$

Comparing those results, we see that, for a given mass, more ions are in NaCl than in $MgCl_2$.

(c) The formula mass for $CrCl_3$ is 158.355 g/mol, and the formula mass for Na_2S is 78.045 g/mol. Because $CrCl_3$ has four ions in its formula and Na_2S has three, which has more ions is not immediately obvious. To determine that, assume a convenient number of grams (10 g), convert the grams to moles of ions for each salt, and compare:

$$10 \text{ g CrCl}_3 \times \frac{1 \text{ mol}}{158.355 \text{ g}} \times \frac{4 \text{ ions}}{\text{CrCl}_3 \text{ formula unit}} = 0.25 \text{ mol ions in CrCl}_3$$

$$10 \text{ g Na}_2\text{S} \times \frac{1 \text{ mol}}{78.045 \text{ g}} \times \frac{3 \text{ ions}}{\text{Na}_2\text{S formula unit}} = 0.38 \text{ mol ions in Na}_2\text{S}$$

Comparing those results, we see that, for a given mass, more ions are in Na_2S than in $CrCl_3$.

Think About It

We can easily imagine in part b that the salt with more ions, not the one with the lowest formula mass, will have more ions. That is not necessarily the case, however, so we must compare them numerically. An example is for MgF_2 (formula weight = 62.3 g/mol, 0.48 mol ions in 10 g) compared with NaCl (formula weight = 58.4 g/mol, 0.34 mol ions in 10 g).

2.81. **Collect and Organize**

Given different amounts of different aluminum-containing substances we are asked to rank them in order of increasing amounts of aluminum.

Analyze

To make the comparisons, we can conveniently convert all the amounts to moles of aluminum.

(a) One mole of Al_2O_3 will contain 2 mol Al.
(b) 76 g of Al_2O_3 will contain 76 g / (101.96 g/mol) × 2 mol Al = 1.49 mol Al.
(c) 27 g of Al will contain 27 g / (26.982 g/mol) = 1.00 mol Al.
(d) 1.1×10^{24} atoms of Al will contain 1.1×10^{24} atoms / (6.022×10^{24} atoms/mol) = 1.8 mol Al.
(e) 1.2 mol of $AlCl_3$ will contain 1.2 mol of Al.

Solve

The order of these substance in increasing amounts of aluminum:

(c) 27 g of Al < (e) 1.2 mol $AlCl_3$ < (b) 76 g Al_2O_3 < (d) 1.1×10^{24} atoms of Al < (a) 1 mol Al_2O_3

Think About It

It's easy to forget in (b) that for every mole of Al_2O_3 calculated from the molar mass, there are two moles of Al.

2.82. **Collect and Organize**

We are to compare the size of a block of 1 mol of aluminum with the size of a block of 1 mol of strontium.

Analyze

Starting with 1 mol of each element, we can obtain the mass by multiplying 1 mol by the element's molar mass (Al = 26.982 g/mol, Sr = 87.62 g/mol). We can find the volume that 1 mol of the element occupies by multiplying the mass by the inverse of the given densities (Al = 1 cm^3/2.70 g, Sr = 1 cm^3/2.64 g). The length of a side of the cube for each element is the cube root of the volume.

Solve

$$1\text{ mol Al} \times \frac{26.982\text{ g}}{1\text{ mol}} \times \frac{1\text{ cm}^3}{2.70\text{ g}} = 9.99\text{ cm}^3$$

$$\text{Length of side of Al cube} = \sqrt[3]{9.99\text{ cm}^3} = 2.15\text{ cm}$$

$$1\text{ mol Sr} \times \frac{87.62\text{ g}}{1\text{ mol}} \times \frac{1\text{ cm}^3}{2.64\text{ g}} = 33.2\text{ cm}^3$$

$$\text{Length of side of Sr cube} = \sqrt[3]{33.2\text{ cm}^3} = 3.21\text{ cm}$$

From those results, the aluminum cube is smaller, with dimensions of 2.15 cm × 2.15 cm × 2.15 cm.

Think About It

The volume of the Sr mole is more than three times greater than that of the Al mole. The volume that 1 mol of a substance takes up is called its *molar volume*. We have calculated that elemental property here for aluminum and strontium.

2.83. **Collect and Organize**

In this question we are asked how mass spectrometry provides information about a molecule.

Analyze

Mass spectrometry plots the mass-to-charge ratio (m/z) versus relative intensity or relative abundance of the peaks.

Solve

In mass spectrometry molecules of the compound are ionized with high-energy electrons to form +1 cations. When a molecule loses one electron, the ion that forms is called the *molecular ion*, M^+. This and other cations produced when the molecule fragments are separated based on their mass-to-charge ratios (m/z). The m/z ratio of a molecular ion with a charge of 1+ corresponds to the molecular mass of the compound. Often, but not always, this is the peak that appears at the highest m/z ratio. Each fragment ion is also counted by the detector giving relative intensities of each fragment ion.

Think About It

Remember that it is the m/z ratio that is determined in mass spectrometry. If a cation has a +2 charge—that is, it loses two electrons in the ionization step—it will appear at an m/z value that is half of its mass.

2.84. **Collect and Organize**

We are to explain the possible origin of the small peak at m/z = 15 u in the mass spectrum of CH_4, which shows a molecular ion peak at m/z = 16 u.

Analyze

Mass spectrometry plots the mass-to-charge ratio (m/z) versus relative intensity or relative abundance of the peaks. Peaks in the mass spectrum may be due to the presence of isotopes of the elements composing the substance or from fragmentation of the molecule.

Solve

The isotopes that might be present in CH_4 (with mass of 16 u) would be present in the mass spectrum at higher *m/z* values (because the next most abundant isotopes for carbon and hydrogen would be ^{13}C and ^{2}H). Therefore, the *m/z* peak at 15 must be due to fragmentation; it must be due to the loss of an H atom from CH_4.

Think About It

If ^{2}H and ^{13}C were more abundant or our mass spectrometer were very sensitive, we would expect to see *m/z* peaks at 17 u ($CH_3{}^2H$ or $^{13}CH_4$), 18 u ($CH_2{}^2H_2$ or $^{13}CH_3{}^2H$), 19 u ($^{13}CH_2{}^2H_2$ or CH^2H_3), 20 u (C^2H_4 or $^{13}CH_3{}^2H$), and 21 u ($^{13}C^2H_4$).

2.85. **Collect and Organize**

We are asked to consider whether the mass spectrum of CO_2 and that of C_3H_8 would show the same molecular ion peak.

Analyze

Mass spectrometry plots the mass-to-charge ratio (*m/z*) versus relative intensity or relative abundance of the peaks. Molecules with the same molecular mass will show the same *m/z* ratio for their molecular ions.

Solve

To the nearest unit, the molecular mass of CO_2 is 44 u and for C_3H_8 the molecular mass also is 44 u. Therefore, both molecules will show the same molecular ion peak at *m/z* 44.

Think About It

That result does not mean that the mass spectra of these two compounds will be the same, however. The two molecules will probably show different fragmentation patterns in their mass spectra.

2.86. **Collect and Organize**

We are asked to consider whether the mass spectrum of CO_2 and that of C_3H_8 would show the same pattern.

Analyze

Mass spectrometry plots the mass-to-charge ratio (*m/z*) versus relative intensity or relative abundance of the peaks. Molecules with the same molecular mass will show the same *m/z* ratio for their molecular ions. In the mass spectrometer, the molecules may also (and often do) fragment in distinctive patterns.

Solve

Because CO_2 and C_3H_8 have very different compositions, we would expect that, despite their having molecular ion peaks at the same *m/z* value due to their identical molecular masses, their fragmentation patterns would be different.

Think About It

The fragmentation pattern of a molecule in the mass spectrum can be used as a fingerprint to identify a molecule.

2.87. **Collect and Organize**

For the explosive materials given we are to calculate the masses of their molecular ion peaks in the mass spectrum.

Analyze

The molecular ion is formed through the loss of one electron from the molecule. The mass of an electron is negligible, so the mass of the molecular ion peak is that of the molecule. We can compute the mass of a molecule using masses of the elements from the periodic table and their ratios in the molecular formula.

Solve

(a) $C_3H_6N_6O_6$: $(3 \times 12\ u) + (6 \times 1\ u) + (6 \times 14\ u) + (6 \times 16\ u) = 222\ u$

(b) $C_4H_8N_8O_8$: $(4 \times 12\ u) + (8 \times 1\ u) + (8 \times 14\ u) + (8 \times 16\ u) = 296\ u$

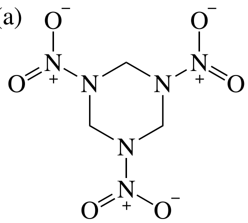
(c) $C_5H_8N_4O_{12}$: $(5 \times 12\ u) + (8 \times 1\ u) + (4 \times 14\ u) + (12 \times 16\ u) = 316\ u$

(d) $C_{14}H_6N_6O_{12}$: $(14 \times 12\ u) + (6 \times 1\ u) + (6 \times 14\ u) + (12 \times 16\ u) = 450\ u$

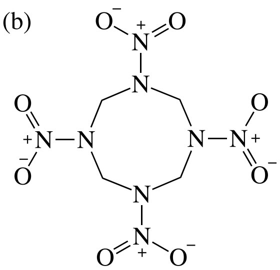
Think About It

These explosives have the following structures. What common features do you see?

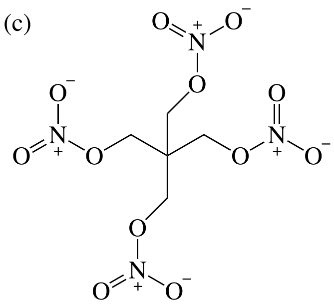
(a)



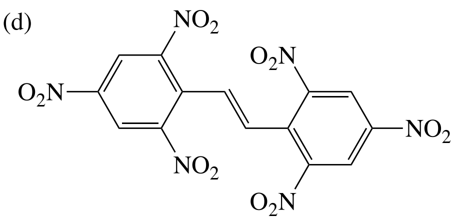
(b)



(c)



(d)



2.88. **Collect and Organize**

For the gases emitted from landfills, CH_4, C_2H_6S, and $C_2H_2Cl_2$, we are to calculate the masses of their molecular ion peaks in the mass spectrum.

Analyze

The molecular ion is formed through the loss of one electron from the molecule. The mass of an electron is negligible, so the mass of the molecular ion peak is that of the molecule. We can compute the mass of a molecule by using masses of the elements from the periodic table and their ratios in the molecular formula.

Solve

CH_4: $(1 \times 12\ u) + (4 \times 1\ u) = 16\ u$
C_2H_6S: $(2 \times 12\ u) + (6 \times 1\ u) + (1 \times 32\ u) = 62\ u$
$C_2H_2Cl_2$: $(2 \times 12\ u) + (2 \times 1\ u) + (2 \times 35.5\ u) = 97\ u$

Think About It

Chlorine has two relatively abundant isotopes, ^{35}Cl (78% abundant) and ^{37}Cl (24% abundant), so we would expect to see three molecular ions peaks for $C_2H_2Cl_2$: one for $C_2H_2{}^{35}Cl_2$, one for $C_2H_2{}^{35}Cl^{37}Cl$, and one for $C_2H_2{}^{37}Cl_2$. At what *m*/*z* values would you see each of those peaks?

2.89. **Collect and Organize**

From the mass spectrum for H_2S shown in Figure P2.89, we are asked to explain how the spectrum shows the sequential loss of H from the molecule. The low-intensity peaks at *m*/*z* ratios above the molecular ion peak suggests the presence of isotopes. We may need the mass numbers and natural abundances of the naturally occurring isotopes of sulfur (Table A3.3).

Analyze

The possible combinations of the sulfur isotopes in H_2S are $H_2{}^{32}S$ (34 u), $H_2{}^{33}S$ (35 u), $H_2{}^{34}S$ (36 u) and $H_2{}^{35}S$ (37 u). Loss of hydrogen from these would decrease *m/z* by 1 u.

Solve

The most intense peak, at 34 u, is the molecular ion $H_2{}^{32}S^+$ (containing the most abundant isotope of sulfur). The peaks at 33 and 32 u are produced when electron bombardment of $H_2{}^{32}S^+$ ions fragments them, forming ions $H^{32}S^+$ and $^{32}S^+$ atoms. The source of the small peak at 36 u is probably the molecular ion of $H_2{}^{34}S$. The even smaller peak at 35 u is probably a combination of mostly $H^{34}S$ and a very little $H_2{}^{33}S$.

Think About It

Fragmentation patterns such as this for H_2S can be important in using mass spectrometry to identify compounds.

2.90. **Collect and Organize**

From the mass spectrum for AsH_3 shown in Figure P2.90 and given that arsenic has only one stable isotope, we are to assign a formula to each peak in the mass spectrum.

Analyze

Arsine, AsH_3, has a molecular mass of 78 u.

Solve

The molecular ion peak AsH_3^+ is assigned to *m*/*z* 78 u. The other three peaks are due to the loss of H atoms with AsH_2^+ at 77 u, AsH^+ at 76 u, and As^+ at 75 u.

Think About It

The AsH^+ peak is the most intense peak, not the molecular ion peak, in this mass spectrum—showing that the molecular ion is not always the most intense peak in the spectrum.

2.91. **Collect and Organize**

Thomson's experiment revealed the electron and its behavior in magnetic and electric fields. In this question, we examine his experiment.

Analyze

Thomson showed that a cathode ray was deflected by a magnetic field in one direction and by an electric field in the other direction. He saw the deflection of the cathode ray when the ray hit a fluorescent plate at the end of his experimental apparatus, as shown in the textbook in Figure 2.2. The cathode ray was deflected by the electrically charged plates as shown. We can imagine the experiment proceeding from no voltage across the charged plates to low voltages and then to higher voltages. From this thought experiment, the ray must be deflected more by an increase in the voltage across the charged plates. Thomson reasoned that the cathode ray was composed of tiny charged particles, which he called "corpuscles."

Solve

(a) Today we call the particles that make up the cathode rays electrons.
(b) The beam of electrons was deflected between the charged plates because they were attracted to the oppositely charged plate and repelled by the negatively charged plate as the beam passed through the electric field. Indeed, in Figure 2.2b we see the beam deflected up toward the (+) plate.
(c) If the polarities of the plates were switched, the electron would still be deflected toward the positively charged plate, which would now be at the bottom of the tube, and the light spot would appear on the bottom of the screen instead of at the top as in Figure 2.2b.

Think About It

That experiment was key to the discovery of subatomic particles, which until then in the atomic theory were not known to exist. It was believed before that time that the atom was the smallest indivisible component of matter.

2.92. **Collect and Organize**

In this problem we are given the masses of the three isotopes of magnesium, from which we can figure out the nuclides (^{24}Mg = 23.9850 u, ^{25}Mg = 24.9858 u, and ^{26}Mg = 25.9826 u). We are also given that the abundance of ^{24}Mg is 78.99%. From that information and the average (weighted) atomic mass of magnesium from the periodic table (24.305 u), we must calculate the abundances of the other two isotopes, ^{25}Mg and ^{26}Mg.

Analyze

The average atomic mass is derived from a weighted average of the isotopes' atomic masses. If x = abundance of ^{25}Mg and y = abundance of ^{26}Mg, the weighted average of magnesium is

$$(0.7899 \times 23.9850) + 24.9858x + 25.9826y = 24.305$$

Because the sum of the abundances of the isotopes must add up to 1.00,

$$0.7899 + x + y = 1.00$$

So

$$x = 1.00 - 0.7899 - y = 0.2101 - y$$

Substituting this expression for x in the weighted average mass equation gives

$$(0.7899 \times 23.9850) + 24.9858(0.2101 - y) + 25.9826y = 24.305$$

Solve

$$18.94575 + 5.249517 - 24.9858y + 25.9826y = 24.305$$
$$0.9968y = 0.1097$$
$$y = 0.1101$$

So

$$x = 0.2101 - 0.1101 = 0.1000$$

The abundance of ^{25}Mg is $x \times 100 = 10.00\%$, and the abundance of ^{26}Mg is $y \times 100 = 11.01\%$.

Think About It

Although the abundances of ^{25}Mg and ^{26}Mg are nearly equal to each other at the end of that calculation, we cannot assume that in setting up the equation. We have to solve the problem algebraically by setting up two equations with two unknowns.

2.93. **Collect and Organize**

Using the periodic table in Figure P2.93, we are to assign atomic numbers to the highlighted elements.

Analyze

The atomic number increases by 1 as we move from hydrogen in the upper left-hand corner of the periodic table to the right and then move down to the next row.

Solve

The atomic numbers for the highlighted elements are as follows: for blue, 9; for yellow, 16; for green, 25; for red, 33; and for lilac, 73.

Think About It

Be careful in identifying the atomic number for the lilac element. The lanthanide elements occur in the periodic table after La and continue through Lu (atomic numbers 58–71)

2.94. **Collect and Organize**

Given the diameter of spherical silver nanoparticles and the number of atoms in one nanoparticle, we are to calculate the number of nanoparticles in 1.00 g.

Analyze

If we divide the mass (1.00 g) of silver by the molar mass of silver and then multiply the result by the Avogadro constant, we will obtain the number of silver atoms in 1.00 g of nanoparticles. If we then divide that result by the number of atoms of silver in one nanoparticle, we will obtain the number of nanoparticles in the 1.00 g sample.

Solve

$$1.00\ \text{g} \times \frac{1\ \text{mol}}{107.87\ \text{g}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}} \times \frac{1\ \text{nanoparticle}}{4.8 \times 10^{7}\ \text{atoms}} = 1.2 \times 10^{14}\ \text{nanoparticles}$$

Think About It

We did not need the information given in this problem pertaining to the diameter of the nanoparticles.

2.95. **Collect and Organize**

For CdS and CdSe quantum dots we are asked to calculate their formula masses, find how many grams of Se are in a sample of CdSe that contains 2.7×10^7 atoms of Cd, and calculate the masses of Cd and S in 4.3×10^{-15} g of CdS.

Analyze

The ratio of Cd to both Se and S in these quantum dots is 1:1. We can use the masses of the elements on the periodic table for Cd, Se, and S to calculate the formula masses of the CdSe and CdS quantum dots. Then we can use the formula mass of CdS with the masses of Cd and S to calculate the masses of Cd and S in a 4.3×10^{-15} g sample of CdS.

Solve

(a) Formula mass of CdS: 112.41 + 32.065 = 144.48 u
Formula mass of CdSe: 112.41 + 78.96 = 191.37 u

(b) $$2.7\times10^{7}\ \text{atoms Cd}\times\frac{1\ \text{atom Se}}{1\ \text{atom Cd}}=2.7\times10^{7}\ \text{atoms Se}$$

(c) $$4.3\times10^{-15}\ \text{g CdS}\times\frac{1\ \text{formula unit CdS}}{144.48\ \text{u}}\times\frac{1\ \text{Cd atom}}{1\ \text{formula unit CdS}}\times\frac{112.41\ \text{u}}{1\ \text{Cd atom}}=3.3\times10^{-15}\ \text{g Cd}$$

$$4.3\times10^{-15}\ \text{g CdS}\times\frac{1\ \text{formula unit CdS}}{144.48\ \text{u}}\times\frac{1\ \text{S atom}}{1\ \text{formula unit CdS}}\times\frac{32.065\ \text{u}}{1\ \text{S atom}}=9.5\times10^{-16}\ \text{g S}$$

Think About It

Which do you think would contain more formula units: 1.0 g of CdSe or 1.0 g of CdS?

2.96. **Collect and Organize**

Given that the average molar mass of air is 28.8 g/mol and that each mole of air had 411.31×10^{-6} mol of CO_2, we are to calculate how many micrograms of CO_2 that sample of air contains.

Analyze

This is a problem in unit analysis. If we multiply the moles of CO_2 per mole of air by the molar mass of air and then multiply that result by the molar mass of CO_2, we will obtain the mass of CO_2 per gram of air.

Solve

$$\frac{411.31\times10^{-6}\ \text{mol CO}_2}{\text{mol air}}\times\frac{\text{mol air}}{28.8\ \text{g air}}\times\frac{44.0\ \text{g CO}_2}{\text{mol CO}_2}\times\frac{1\ \mu\text{g CO}_2}{1\times10^{-6}\ \text{g CO}_2}=628\ \mu\text{g CO}_2/\text{g air}$$

Think About It

Always label your values with units and make sure that they cancel correctly to arrive at your final answer.

2.97. **Collect and Organize**

From the mass spectrum of cocaine shown in Figure P2.97, we are to determine the molar mass.

Analyze

The molar mass can be determined from a mass spectrum from the molecular ion peak. The molecular ion peak is (usually) the peak at the largest *m*/*z* value.

Solve

The largest *m*/*z* value is at 303 u. Therefore, the molar mass of cocaine is 303 g/mol.

Think About It

The molecular ion peak is not the most intense peak in this mass spectrum.

2.98. **Collect and Organize**

From the mass spectrum shown in Figure P2.98, we are to identify the molecular ion peak and show that it is consistent with the formula $C_{21}H_{28}O_2$.

Analyze

The molecular ion peak is (usually) the peak at the largest *m*/*z* value. We can compute the expected mass of the compound by using values from the periodic table for C, H, and O.

Solve

The molecular ion peak is at 312 u, and that is consistent with the mass of $C_{21}H_{28}O_2$, which is (21 × 12 u) + (28 × 1 u) + (2 × 16 u) = 312 u.

Think About It

The molecular ion peak is not the most intense peak, but the largest mass peak, in a mass spectrum.

2.99. **Collect and Organize**

To determine the number of moles and atoms of carbon in the Hope Diamond, we have to first convert its given mass of 45.52 carats to mass in grams.

Analyze

To convert the mass of the diamond to grams, we use the relationship 1 carat = 200 mg. The moles of carbon atoms is equal to the mass of the diamond divided by the molar mass of carbon (12.011 g/mol); the number of carbon atoms in the diamond is the number of moles of carbon multiplied by the Avogadro constant.

Solve

(a) The mass of the diamond in grams is

$$45.52 \text{ carats} \times \frac{200 \text{ mg}}{1 \text{ carat}} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 9.104 \text{ g}$$

The number of moles of carbon atoms in the diamond is

$$9.104 \text{ g} \times \frac{1 \text{ mol}}{12.011 \text{ g}} = 0.7580 \text{ mol of C}$$

(b) The number of carbon atoms in the diamond is

$$0.7580 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 4.565 \times 10^{23} \text{ atoms of C}$$

Think About It

Diamond is the hardest natural substance and is an electrical insulator while being an excellent conductor of heat.

2.100. **Collect and Organize**

Given the molecular formulas for three compounds that compose the minerals called feldspars, we are to rank them in order of increasing percentage Al by mass.

Analyze

The percent by mass for each of the compounds is determined by dividing the mass of the aluminum in the formula (mol of Al × 26.982 u) by the mass of the formula unit of the compounds, then multiplying by 100 for the percent. Using the periodic table, we can calculate the masses of the compounds:

(a) $NaAlSi_3O_8$ = 22.990 u + 26.982 u + 3(28.086 u) + 8(15.999 u) = 262.22 u
(b) $KAlSi_3O_8$ = 39.098 u + 26.982 u + 3(28.086 u) + 8(15.999 u) =278.33 u
(c) $CaAl_2Si_2O_8$= 40.078 u + 2(26.982 u) + 2(28.086 u) + 8(15.999 u) = 278.21 u.

Solve

(a) $NaAlSi_3O_8$

$$\frac{26.982 \text{ u}}{262.22 \text{ u}} \times 100 = 10.290\%$$

(b) $KAlSi_3O_8$

$$\frac{26.982\ \text{u}}{278.33\ \text{u}} \times 100 = 9.6942\%$$

(c) $CaAl_2Si_2O_8$

$$\frac{2 \times 26.982\ \text{u}}{278.21\ \text{u}} \times 100 = 19.397\%$$

In order of increasing percent by mass of aluminum: (b) $KAlSi_3O_8$ < (a) $NaAlSi_3O_8$ < (c) $CaAl_2Si_2O_8$

Think About It

Remember that in multiplication or division we use the rule of the number of the least significant figures; therefore, the percentages here have 5 significant figures.