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 it's 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 the neutrons, the 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 element, 118, 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, and a halogen.

Analyze

Transition metals occupy the middle of the periodic table and reside in the *d*-block, alkali metals occupy the first column of the periodic table in the *s*-block, and halogens occupy the second-to-last column on the right, 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).

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 this question we need to correlate properties of the elements with their positions in the periodic table. We need to use the definition of *inert* and know the general regions of the periodic table that have gaseous, metallic, and nonmetallic elements.

Analyze

Inert means that the chemical species does not (or does not readily) combine with other species. Here we are looking for elements that are relatively unreactive as a group; these are the noble gases. On the periodic table, metallic elements tend to be on the left-hand side; nonmetallic elements are on the right-hand side. The gases tend to be in groups 18 (noble gases) and 17 (the halogens, fluorine and chlorine) along with oxygen, nitrogen, and hydrogen. Of these gases, the noble gases are monatomic, but the others are all diatomic (N_2, O_2, H_2, F_2, Cl_2).

Solve

In the periodic table shown in Figure P2.4, elements Na, Ne, Cl, Au, and Lr are highlighted.

(a) Chlorine (Cl_2, lilac) is a diatomic gas at room temperature. Neon (Ne, red) is a monatomic gas at room temperature. Both are nonmetals.
(b) Neon (Ne, red) is a chemically inert gas.
(c) Sodium (Na, dark blue), gold (Au, green), and lawrencium (Lr, light orange) are all metals.

Think About It

As metals, gold and lawrencium differ in their properties compared with sodium (which is highly reactive). Gold can be found as an element in nature and is not very reactive, and lawrencium is radioactive and synthesized in small amounts; therefore, its chemistry is relatively unexplored.

2.7. **Collect and Organize**

Using Figure P2.7 and our knowledge of the charges and the masses of alpha and beta 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 betas 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 must represent a neutral particle because it is not deflected by the 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 green arrow in Figure P2.7.

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

Solve
The green arrow in Figure P2.7 represents electrons as negatively charged particles.

Think About It
Other subatomic particles represented in the diagram are protons (red arrow) and 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 amu and another grouping around 49 amu.

Solve
The molecular ion peaks showing up at 84, 86, and 88 amu 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 amu is due to the presence of two atoms of ^{35}Cl ($CH_2{}^{35}Cl_2$), the m/z peak at 86 amu 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 amu is due to the presence of two atoms of ^{37}Cl ($CH_2{}^{37}Cl_2$). Also, the m/z peak grouping around 50 amu is consistent with 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.

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

Solve
The most abundant isotope has the highest relative intensity; for krypton that is an *m*/*z* of 84 amu. 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.

Think About It
The other isotopes have 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 2.10 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), 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.

Analyze

The plum-pudding model of the atom viewed the electrons as small particles in a diffuse, positively charged "pudding." In Rutherford's experiment, most of the 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 "pudding."

Solve

Had the plum-pudding model been valid, the alpha 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 (Figure A2.14).

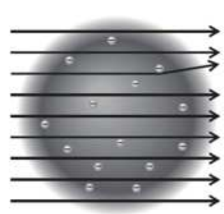


Figure A2.14.

Think About It
Because the positive charge in the atom was viewed as diffuse, no alpha particles would have bounced back toward their source in an experiment confirming the plum-pudding model. The key result of Rutherford's experiment was that a few alpha particles came directly back toward the alpha 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 ray through a magnetic field, and he discovered that the ray was deflected. A magnetic field would not deflect pure-energy "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) are the traditional (that is, not LCD) television and computer screens.

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 two 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 its products of further decays) decays by α emission. Alpha particles, composed of two protons and two neutrons, easily pick up electrons from their environment to become helium.

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 different 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 thickness 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 how the result would have been different if Rutherford's gold atoms in the foil had absorbed the α particles.

Analyze

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

Solve

We would see fewer α particles either passing directly through the foil or being deflected at large angles. Absorption of an α particle by a gold atom would have transmuted the gold atom into the heavier element, thallium.

Think About It

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

$$^{197}Au + \alpha \rightarrow {}^{201}Tl$$

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 in the Rutherford experiment.

Analyze

Here we want to look at how gold differs from silver and aluminum.

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 relate the mass number and the atomic number to the number of protons and neutrons in the nucleus of an atom.

Analyze

Atomic number is the count of the protons in the nucleus; the mass number is the number of protons and neutrons combined.

Solve

To get the atomic number for a nuclide, we count the protons in the nucleus. To arrive at the mass number, we sum the number of protons with the number of neutrons in the nucleus. In other words, the mass number minus the atomic number is the number of neutrons in the nucleus of the nuclide.

Think About It

The atomic number defines the element on the periodic table. The mass number gives us information about the isotope(s) of that element.

2.23. **Collect and Organize**

Given that most stable nuclides have at least the same numbers of 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 is the exception.

Think About It

Hydrogen, however, can have one (for deuterium) and even two (for tritium) neutrons in its nucleus.

2.24. **Collect and Organize**

We are to explain the redundancy in writing a nuclide with the symbol ${}^{A}_{Z}\mathrm{X}$.

Analyze

A gives the sum of protons and neutrons in the nucleus, Z gives the number of protons in the nucleus, and X identifies the element by its two-letter symbol on the periodic table.

Solve
Both Z and X describe the element on the periodic table, so only one is necessary.

Think About It
Often the nuclides are written with Z missing, as in ^{235}U.

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 for the isotope, 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 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 for the isotope, 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 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 ratio for group 15 nuclei ^{14}N, ^{31}P, ^{75}As, and ^{121}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 ^{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 for the element symbol.

Analyze

An isotope is given by the symbol ${}^{A}_{Z}\mathrm{X}$, where X is the element 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 proton and electrons.

Solve

Symbol	${}^{23}\mathrm{Na}$	${}^{89}\mathrm{Y}$	${}^{118}\mathrm{Sn}$	${}^{197}\mathrm{Au}$
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 for the element symbol.

Analyze

An isotope is given by the symbol ${}^{A}_{Z}\mathrm{X}$, where X is the element 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 proton and electrons.

Solve

Symbol	${}^{27}\mathrm{Al}$	${}^{98}\mathrm{Mo}$	${}^{143}\mathrm{Nd}$	${}^{238}\mathrm{U}$
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 is given by the symbol ${}^{A}_{Z}\mathrm{X}^{n}$, where X is the element 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 by adding the protons to the neutrons in the nucleus for the isotope. We can determine the number of neutrons or protons in the nucleus for the isotopes by subtracting Z (number of protons) or the number of neutrons from A (mass number), respectively. 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 is given by the symbol $^{A}_{Z}X^{n}$, where X is the element 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 by adding the protons to the neutrons in the nucleus for the isotope. We can determine the number of neutrons or protons in the nucleus for the isotopes by subtracting Z (number of protons) or the number of neutrons from A (mass number), respectively. 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 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
This classification was based on the chemical behavior of the elements, which was Mendeleev's brilliant insight.

2.34. **Collect and Organize**
Knowing that Mendeleev labeled his groups on the right 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.

Solve
The noble gases were not discovered until after Mendeleev put together his periodic table. He also could not have predicted the existence of the noble gases at the time since (a) none of them was 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 to describe the periodic trends (across a row and down a group) in charges on monatomic ions.

Analyze
We can use Figure 2.10 in the textbook to see the trends described in the question.

Solve
As we proceed across a row in the periodic table and the group number increases, we first start off with a low 1+ charge for the alkali metals but then increase more positively by 1 through the alkaline earth, boron, and carbon groups. Then the trend shifts to highly negative charges, as in 3– for nitrogen and phosphorus and then reducing one unit through oxygen (2–) and fluorine (1–) to zero for neon.

As we proceed down a group on the periodic table, the charge on the monatomic ion stays the same except for the heavy group 13, 14, and (not shown in Figure 2.10) 15 elements, which show two charges: one like the rest of its lighter group members and the other two units less positive.

Think About It

The similar charges for elements in the same group (column) of the periodic table results in their similar chemistries.

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 suffocation when inhaled.

2.39. **Collect and Organize**

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

Analyze

In examining the clues in parts a–c, we can guess that these will all be transition metals.

Solve

(a) The element in the fifth row of the periodic table (Rb–Xe) in group 10 is palladium (Pd).
(b) The element to the left of Pd in group 9 is rhodium (Rh).
(c) The element below Pd in group 10 is platinum (Pt).

Think About It

These metals are generally as expensive as gold. Platinum is about \$1400/oz and gold is about \$1200/oz. Rhodium and palladium are less expensive, at about \$1000/oz and \$700/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 semimetal elements begin at silicon and move into the nonmetals 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
As a semimetal, silicon has some, but not all, the properties of a metal but has a structure more like that of a nonmetal.

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. Those 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 semimetal described.

Think About It
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, all the masses in the periodic table for the elements are calculated weighted averages.

2.44. **Collect and Organize**

In this question we relate the percentage of natural abundance to the average atomic mass of an element.

Analyze

The natural abundance of an isotope of an element indicates how common that isotope is in nature (on Earth).

Solve

Isotopic abundances are used to compute the weighted average atomic mass. The known isotopic abundances must be considered since the dominance of one isotope will contribute most to the average mass. To obtain the molar mass for an element on the periodic table, we multiply the percent abundance of each naturally occurring isotope by the mass of that isotope and then sum those values.

Think About It

The high abundance of one isotope over another indicates its higher nuclear stability.

2.45. **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?

2.46. **Collect and Organize**

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

Analyze

For a cation, the actual mass would be the atomic mass minus the mass of the 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 the atom, so the mass of the ions is exceedingly close to the mass of the neutral atoms. But more important, 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 $MgCl_2$?

2.47. **Collect and Organize**

In this question we are asked to explain the observation that the average atomic mass of platinum is 195.08 amu, 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 amu 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 amu.

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

2.48. **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 amu 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.

2.49. **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 amu.

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

Think About It
Indeed, ^{40}Ar is 99.6% abundant.

2.50. **Collect and Organize**
Given the information that manganese has only one stable isotope, we can use the mass from the periodic table to determine the number of neutrons in its nucleus.

Analyze
Manganese has an atomic mass of 54.938 amu (which is very close to 55) and an atomic number of 25.

Solve
The number of neutrons in the manganese nucleus is the mass number (55) minus the atomic number (25), so manganese has 30 neutrons.

Think About It

Other elements that have 100% natural abundance include ^{19}F, ^{27}Al, and ^{31}P.

2.51. **Collect and Organize**

We have to consider the concept of weighted average atomic mass to answer this question.

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 are 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 amu. The actual average mass (10.811 amu) 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 amu. The actual average mass (6.941 amu) 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 amu. The actual average mass (14.007 amu) 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.

2.52. **Collect and Organize**

We have to consider the concept of weighted average atomic mass to answer this question.

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 are 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 ^{85}Rb and ^{87}Rb would be 86 amu. The actual average mass (85.468 amu) is less than this; therefore, ^{85}Rb is more abundant.
(b) The simple average atomic mass for ^{69}Ga and ^{71}Ga would be 70 amu. The actual average mass (69.723 amu) is less than this; therefore, ^{69}Ga is more abundant.
(c) The simple average atomic mass for ^{50}V and ^{51}V would be 50.5 amu. The actual average mass (50.942 amu) is greater than this; therefore, ^{51}V is more abundant.

Think About It

This is a quick question to answer for elements such as rubidium, gallium, and vanadium 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.

2.53. **Collect and Organize**

In this question we are given the masses and abundances of the naturally occurring isotopes of copper. From that information, we can calculate the average atomic mass of copper.

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 copper

$$m_{\mathrm{Cu}} = (0.6917 \times 62.9296 \text{ amu}) + (0.3083 \times 64.9278 \text{ amu}) = 63.55 \text{ amu}$$

Think About It

Because copper-63 is more abundant than copper-65, we expect that the average atomic mass for copper would be below the simple average of 64.

2.54. **Collect and Organize**

In this question we are given the masses and abundances of the four naturally occurring isotopes of sulfur. From that information, we can calculate the average atomic mass of sulfur.

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 sulfur

$$m_{\mathrm{S}} = (0.9504 \times 31.9721 \text{ amu}) + (0.0075 \times 32.9715 \text{ amu}) + (0.0420 \times 33.9679 \text{ amu}) + (0.0001 \times 35.9671 \text{ amu}) = 32.1 \text{ amu}$$

Think About It

Because sulfur-32 is about 95% abundant and the other isotopes are present in low abundances, we would expect the average atomic mass to be close to the mass of sulfur-32.

2.55. **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

$$m_{\mathrm{Mg}} = (0.7870 \times 23.9850 \text{ amu}) + (0.1013 \times 24.9858 \text{ amu}) + (0.1117 \times 25.9826 \text{ amu}) = 24.31 \text{ amu}$$

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.

2.56. **Collect and Organize**

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

Analyze

The weighted atomic mass is computed according to the formula

$$m_x = a_1 m_1 + a_2 m_2 + a_3 m_3 + \cdots$$

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

$$m_{\mathrm{Sr}} = (0.0056 \times 83.9134 \text{ amu}) + (0.0986 \times 85.9094 \text{ amu}) + (0.0700 \times 86.9089 \text{ amu}) + (0.8258 \times 87.9056 \text{ amu}) = 87.62 \text{ amu}$$

The atomic mass of Sr in the periodic table is 87.62 amu. These values match.

Think About It

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

2.57. **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 ^{48}Ti isotope.

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,

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

Solve

$$47.867 \text{ amu} = (0.0825 \times 45.9526) + (0.0744 \times 46.9518) + (0.7372 \times m_{^{48}\mathrm{Ti}}) + (0.0541 \times 48.94787) + (0.0518 \times 49.94479)$$

$$m_{^{48}\mathrm{Ti}} = 47.948 \text{ amu}$$

Think About It

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

2.58. **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,

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

Solve

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

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

Think About It

That answer makes sense because the exact mass of ^{92}Ti should be close to 92 amu.

2.59. **Collect and Organize**

From the formulas for three ionic compounds, CaF_2, Na_2S, and Cr_2O_3, 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) CaF_2: 40.078 amu + 2(18.998 amu) = 78.074 amu
(b) Na_2S: 2(22.990 amu) + 32.065 amu = 78.045 amu
(c) Cr_2O_3: 2(51.996 amu) + 3(15.999 amu) = 151.989 amu

Think About It

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

2.60. **Collect and Organize**

From the formulas for three ionic compounds, KCl, MgO, and Al_2O_3, 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) KCl: 39.098 amu + 35.453 amu = 74.551 amu
(b) MgO: 24.305 amu + 15.999 amu = 40.304 amu
(c) Al_2O_3: 2(26.982 amu) + 3(15.999 amu) = 101.961 amu

Think About It

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

2.61. **Collect and Organize**

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

Analyze

The subscript in each formula after the C atom is the number of carbons in the molecule.

Solve

(a) 1
(b) 3
(c) 6
(d) 6

Think About It

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

2.62. **Collect and Organize**

For each given molecular formula in P2.61, we are to determine the number of hydrogen atoms in each.

Analyze

The subscript in each formula after the H atom is the number of hydrogens in the molecule.

Solve

(a) 4
(b) 8
(c) 6
(d) 12

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.

2.63. **Collect and Organize**

For a list of five compounds, we are to determine their molecular masses and then rank them in order of increasing mass.

Analyze

When we use the masses on the periodic table to calculate the molecular masses, we obtain the following:
(a) CO = 28.01 amu
(b) Cl_2 = 70.91 amu
(c) CO_2 = 44.01 amu
(d) NH_3 = 17.03 amu
(e) CH_4 = 16.04 amu

Solve

In order of increasing molecular mass: (e) CH_4 < (d) NH_3 < (a) CO < (c) CO_2 < (b) Cl_2.

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 than four significant digits.

2.64. **Collect and Organize**

For a list of five compounds, we are to determine their molecular masses and then rank them in order of decreasing mass.

Analyze

When we use the masses on the periodic table to calculate the molecular masses, we obtain the following:
(a) H_2 = 2.02 amu
(b) Br_2 = 159.8 amu
(c) NO_2 = 46.01 amu
(d) C_2H_2 = 26.04 amu
(e) BF_3 = 67.81 amu

Solve

In order of decreasing molecular mass: (b) Br_2 > (e) BF_3 > (c) NO_2 > (d) C_2H_2 > (a) H_2.

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 than four significant digits.

2.65. **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.

2.66. **Collect and Organize**
For an ionic compound, we are asked to distinguish between its molar mass and its formula mass.

Analyze
The molar mass is the mass of 1 mol of the particles that compose a substance, and the formula mass is the mass of one formula unit of an ionic compound.

Solve
Both terms give us the mass of 1 mol of a substance. The distinction lies in the fact that the formula mass is reserved for the mass of a mole of a substance that does not have a defined molecular structure, such as ionic compounds held together by the attraction between the negative anions and the positive cations.

Think About It
For a molecular compound, the formula mass is equal to the molecular mass. Those terms are often used interchangeably.

2.67. **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. Substances (including isotopes) with higher molar masses will have fewer atoms than those with lower molar masses.

Solve
No, 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.

2.68. **Collect and Organize**
Given the natural abundance of isotopes of an element in percent by mass, we are to consider whether the same percent values would be true for mole percent of the isotopes.

Analyze
We compute moles of a substance by dividing the mass by the molar mass.

Solve
Let's try this out. Boron has two naturally occurring isotopes, ^{10}B (19.9% abundant) and ^{11}B (80.1% abundant). If we have 100 g of naturally occurring boron, let's calculate the mole percent from this mass percent. This would mean that 100 g would have 19.9 g of ^{10}B and 80.1 g of ^{11}B. From that and the molar masses of those isotopes we can determine the number of moles of each.

$$\text{mol } ^{10}\text{B} = 19.9 \text{ g} \times \frac{\text{mol}}{10.01 \text{ g}} = 1.988 \text{ mol}$$

$$\text{mol } ^{10}\text{B} = 80.1 \text{ g} \times \frac{\text{mol}}{11.01 \text{ g}} = 7.275 \text{ mol}$$

The total moles of boron in 100 g is 1.988 + 7.275 = 9.263 mol, and so the percentage by mole of each isotope in this sample is

$$\%\text{ mol } ^{10}\text{B} = \frac{1.988}{9.263} \times 100 = 21.46\%$$

$$\%\text{ mol } ^{11}\text{B} = \frac{7.275}{9.263} \times 100 = 78.54\%$$

The mole percent of the isotopes, therefore, is not the same as the mass percent.

Think About It

The difference we calculated between the mole percent and the mass percent may not seem like much, but it is significant.

2.69. **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 Avogadro's number.

Solve

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

(b) $\dfrac{4.2 \times 10^{13} \text{ molecules of CH}_4}{6.022 \times 10^{23} \text{ molecules/mol}} = 7.0 \times 10^{-11} \text{ mol CH}_4$

(c) $\dfrac{2.5 \times 10^{12} \text{ molecules of O}_3}{6.022 \times 10^{23} \text{ molecules/mol}} = 4.2 \times 10^{-12} \text{ mol O}_3$

(d) $\dfrac{4.9 \times 10^{9} \text{ molecules of NO}_2}{6.022 \times 10^{23} \text{ molecules/mol}} = 8.1 \times 10^{-15} \text{ mol NO}_2$

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.

2.70. **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 Avogadro's number.

Solve

(a) $\dfrac{1.4 \times 10^{13} \text{ molecules of H}_2}{6.022 \times 10^{23} \text{ molecules/mol}} = 2.3 \times 10^{-11} \text{ mol H}_2$

(b) $\dfrac{1.5 \times 10^{14} \text{ atoms of He}}{6.022 \times 10^{23} \text{ atoms/mol}} = 2.5 \times 10^{-10} \text{ mol He}$

(c) $\dfrac{7.7 \times 10^{12} \text{ molecules of N}_2\text{O}}{6.022 \times 10^{23} \text{ molecules/mol}} = 1.3 \times 10^{-11} \text{ mol N}_2\text{O}$

(d) $\dfrac{3.0 \times 10^{12} \text{ molecules of CO}}{6.022 \times 10^{23} \text{ molecules/mol}} = 5.0 \times 10^{-12} \text{ mol CO}$

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.

2.71. **Collect and Organize**

From the chemical formulas for various iron compounds with oxygen, we are asked to determine how many moles of iron 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 iron is in the compound's chemical formula, then 1 mol of iron is in 1 mol of the compound. Likewise, if three atoms of iron are in the chemical formula, 3 mol of iron is present in 1 mol of the substance.

Solve

(a) One atom of iron is in FeO; therefore, 1 mol of FeO contains 1 mol of iron.
(b) Two atoms of iron are in Fe_2O_3; therefore, 1 mol of Fe_2O_3 contains 2 mol of iron.
(c) One atom of iron is in $Fe(OH)_3$; therefore, 1 mol of $Fe(OH)_3$ contains 1 mol of iron.
(d) Three atoms of iron are in Fe_3O_4; therefore, 1 mol of Fe_3O_4 contains 3 mol of iron.

Think About It

The parentheses used in $Fe(OH)_3$ show that three OH units are in that compound. If the question had asked how many moles of oxygen were present in 1 mol of that substance, the answer would be 3 mol of oxygen.

2.72. **Collect and Organize**

From the chemical formulas for various sodium salts, we are asked to determine how many moles of Na^+ 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 sodium is in the compound's chemical formula, then 1 mol of Na^+ is in 1 mol of the compound. Likewise, if two atoms of sodium are in the chemical formula, 2 mol of Na^+ is present in 1 mol of the substance.

Solve

(a) One atom of sodium is in $NaCl$; therefore, 1 mol of $NaCl$ contains 1 mol of Na^+.
(b) Two atoms of sodium are in Na_2SO_4; therefore, 1 mol of Na_2SO_4 contains 2 mol of Na^+.
(c) Three atoms of sodium are in Na_3PO_4; therefore, 1 mol of Na_3PO_4 contains 3 mol of Na^+.
(d) One atom of sodium is in $NaNO_3$; therefore, 1 mol of $NaNO_3$ contains 1 mol of Na^+.

Think About It

The number of sodium ions appearing in those chemical formulas balances the charge on the anions in the salts. For example, three sodium cations are needed to produce the neutral salt of PO_4^{3-}.

2.73. **Collect and Organize**

We are to calculate the mass of a given number of moles of magnesium carbonate.

Analyze

To convert from moles to mass, multiply the number of moles by the molar mass of the substance. The molar mass of $MgCO_3$ is 24.30 + 12.01 + 3(16.00) = 84.31 g/mol.

Solve

$$0.122 \text{ mol MgCO}_3 \times \frac{84.31 \text{ g}}{\text{mol}} = 10.3 \text{ g}$$

Think About It

Moles in chemistry are like a common currency in exchanging money. From moles we can calculate mass; from mass we can calculate moles.

2.74. **Collect and Organize**

We are asked to determine the volume that 1.00 mol of benzene, a liquid, occupies at 20˚C.

Analyze

To answer this question, we have to first find the mass of 1.00 mol of benzene (C_6H_6). Once we have the mass, we can use the density at 20°C to find the volume. Benzene's molar mass is 6(12.01) + 6(1.01) = 78.12 g/mol.

Solve

$$1.00 \text{ mol } C_6H_6 \times \frac{78.12 \text{ g}}{\text{mol}} \times \frac{1 \text{ mL}}{0.879 \text{ g}} = 88.9 \text{ mL}$$

Think About It

The mass of 1.00 mol of benzene weighs 78.10 g, but the volume it takes up is more than 78 mL because its density is less than that of water (1.00 g/mL).

2.75. **Collect and Organize**

In this exercise we convert from the moles of titanium contained in a substance to the number of atoms present.

Analyze

For each substance, we need to take into account the number of moles of titanium *atoms* present in *1 mol* of the substance. For 0.125 mol of substance, then, a substance that contains two atoms of titanium in its formula contains 0.125 × 2 = 0.250 mol of titanium. We can then use Avogadro's number to convert the moles of titanium to the number of atoms present in the sample.

Solve

(a) Ilmenite, $FeTiO_3$, contains one atom of Ti per formula unit, so 0.125 mol of ilmenite contains 0.125 mol of Ti.

$$0.125 \text{ mol Ti} \times \frac{6.022 \times 10^{23} \text{ Ti atoms}}{1 \text{ mol}} = 7.53 \times 10^{22} \text{ Ti atoms}$$

(b) The formula for titanium(IV) chloride is $TiCl_4$. This formula contains only one Ti atom per formula unit as well, so the answer is identical to that calculated in (a).

$$0.125 \text{ mol Ti} \times \frac{6.022 \times 10^{23} \text{ Ti atoms}}{1 \text{ mol}} = 7.53 \times 10^{22} \text{ Ti atoms}$$

(c) Ti_2O_3 contains two titanium atoms in its formula, so 0.125 mol of Ti_2O_3 contains 0.125 × 2 = 0.250 mol of titanium.

$$0.250 \text{ mol Ti} \times \frac{6.022 \times 10^{23} \text{ Ti atoms}}{1 \text{ mol}} = 1.51 \times 10^{23} \text{ Ti atoms}$$

(d) Ti_3O_5 contains three titanium atoms in its formula, so 0.125 mol of Ti_3O_5 contains 0.125 × 3 = 0.375 mol of titanium.

$$0.375 \text{ mol Ti} \times \frac{6.022 \times 10^{23} \text{ Ti atoms}}{1 \text{ mol}} = 2.26 \times 10^{23} \text{ Ti atoms}$$

Think About It

The number of atoms of titanium in 0.125 mol of each compound reflects the number of atoms of Ti in the chemical formula. Ti_2O_3 has twice the number of Ti atoms, and Ti_3O_5 has three times the number of Ti atoms, compared with the number of Ti atoms in the same number of moles of $FeTiO_3$ and $TiCl_4$.

2.76. **Collect and Organize**

In this exercise we convert from the moles of iron contained in a substance to the number of atoms present.

Analyze

For each substance we need to take into account the number of moles of iron *atoms* present in *1 mol* of the substance. For 2.5 mol of substance, then, a substance that contains two atoms of iron in its formula contains 2.5 × 2 = 5.0 mol of iron. We can then use Avogadro's number to convert the moles of iron to the number of atoms present in the sample.

Solve

(a) Wolframite, $FeWO_4$, contains one atom of Fe per formula unit, so 2.5 mol of wolframite contains 2.5 mol of Fe.

$$2.5 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol}} = 1.5 \times 10^{24} \text{ Fe atoms}$$

(b) The formula for pyrite, FeS_2, contains only one iron atom per formula unit as well, so the answer is identical to that calculated in part a.

$$2.5 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol}} = 1.5 \times 10^{24} \text{ Fe atoms}$$

(c) Magnetite, Fe_3O_4, contains three iron atoms in its formula, so 2.5 mol of Fe_3O_4 contains $2.5 \times 3 = 7.5$ mol of iron.

$$7.5 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol}} = 4.5 \times 10^{24} \text{ Fe atoms}$$

(d) Hematite, Fe_2O_3, contains two iron atoms in its formula, so 2.5 mol of hematite contains $2.5 \times 2 = 5.0$ mol of iron.

$$5.0 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol}} = 3.0 \times 10^{24} \text{ Fe atoms}$$

Think About It

The ratio of iron atoms in wolframite to pyrite to magnetite to hematite of 1:1:3:2 reflects the number of iron atoms in their formulas.

2.77. **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 Al_2O_3 contains 3 mol of oxygen, and 1 mol of Fe_2O_3 also contains 3 mol of oxygen. Those compounds contain the same number of moles of oxygen.
(b) One mole of SiO_2 contains 2 mol of oxygen, and 1 mol of N_2O_4 contains 4 mol of oxygen. Therefore, N_2O_4 contains more moles of oxygen (twice as much).
(c) Three moles of CO contains 3 mol of oxygen, and 2 mol of CO_2 contains 4 mol of oxygen. Therefore, the 2 mol of CO_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 concluded wrongly that 3 mol of CO contains more moles of oxygen than 2 mol of CO_2.

2.78. **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) Two moles of N_2O contains 2 mol of oxygen, and 1 mol of N_2O_5 contains 5 mol of oxygen. Therefore, N_2O_5 contains the most moles of oxygen.
(b) One mole of NO contains 1 mol of oxygen, and 1 mol of calcium nitrate [$Ca(NO_3)_2$] contains 6 mol of oxygen. Therefore, calcium nitrate contains six times more moles of oxygen than NO.
(c) Two moles of NO_2 contains 4 mol of oxygen, and 1 mol of sodium nitrite ($NaNO_2$) contains 2 mol of oxygen. NO_2 contains more moles of oxygen.

Think About It

We cannot decide which substance has more moles of oxygen by comparing only the amounts of substances present. If that were the case, we would have concluded wrongly that 2 mol of N_2O contains more moles of oxygen than 1 mol of N_2O_5.

2.79. **Collect and Organize**

For each aluminosilicate, we are given the chemical formula. From that formula we are asked to deduce the number of moles of aluminum in 1.50 mol of each substance.

Analyze

The number of moles of aluminum in 1 mol of each substance is reflected in its chemical formula. We need next to take into account that we are starting with 1.50 mol of each substance.

Solve

(a) Each mole of pyrophyllite, $Al_2Si_4O_{10}(OH)_2$, contains 2 mol of Al atoms. Therefore, 1.50 mol of pyrophyllite contains 1.50 mol × 2 = 3.00 mol of Al.
(b) Each mole of mica, $KAl_3Si_3O_{10}(OH)_2$, contains 3 mol of Al. Therefore, 1.50 mol of mica contains 1.50 mol × 3 = 4.50 mol of Al.
(c) Each mole of albite, $NaAlSi_3O_8$, contains 1 mol of Al. Therefore, 1.50 mol of albite contains 1.50 mol of Al.

Think About It

Those minerals could all be distinguished by analyzing the amount of aluminum present in the same number of moles of each substance.

2.80. **Collect and Organize**

For each uranium mineral we are given the chemical formula. From that formula we are asked to deduce the number of moles of uranium in 1 mol of each substance.

Analyze

The number of moles of uranium in 1 mol of each substance is reflected in its chemical formula.

Solve

(a) Each mole of carnotite, $K_2(UO_2)_2(VO_4)_2$, contains 2 mol of U atoms [from the $(UO_2)_2$ unit].
(b) Each mole of uranophane, $CaU_2Si_2O_{11}$, contains 2 mol of U.
(c) Each mole of autunite, $Ca(UO_2)_2(PO_4)_2$, contains 2 mol of U.

Think About It

All those minerals have the same number of moles of uranium in 1 mol of the mineral.

2.81. **Collect and Organize**

This exercise has us compute the molar mass of various molecular compounds of oxygen.

Analyze

We can find the molar mass of each compound 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 compounds SO_2, O_3, and CO_2 have three atoms in their chemical formula, but each has a different molar mass.

2.82. **Collect and Organize**

This exercise has us compute the molar mass 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) lmenite, $FeTiO_3$: 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.

2.83. **Collect and Organize**

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

Analyze

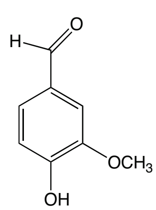
We can find the molar mass of each flavoring by adding the molar mass of each element from the periodic table, taking into account the number of moles of each atom present in 1 mol of the flavoring. Each flavoring contains only carbon (12.01 g/mol), hydrogen (1.01 g/mol), and oxygen (16.00 g/mol).

Solve

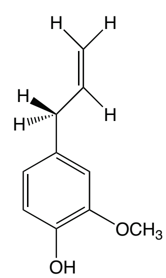
(a) Vanillin, $C_8H_8O_3$: 8(12.011) + 8(1.0079) + 3(15.999) = 152.148 g/mol
(b) Oil of cloves, $C_{10}H_{12}O_2$: 10(12.011) + 12(1.0079) + 2(15.999) = 164.203 g/mol
(c) Anise oil, $C_{10}H_{12}O$: 10(12.011) + 12(1.0079) + 15.999 = 148.204 g/mol
(d) Oil of cinnamon, C_9H_8O: 9(12.011) + 8(1.0079) +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

Think About It

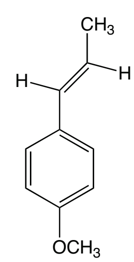
Each flavoring has a distinctive odor and flavor due in part to its different chemical formula. Another factor, however, in differentiating these flavorings is their chemical structure, or the arrangement in which the atoms are attached, as shown by the structures of those flavorings:



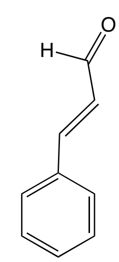
Vanillin



Oil of cloves



Anise oil



Oil of cinnamon

2.84. **Collect and Organize**

This exercise has us compute the molar mass 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).

2.85. **Collect and Organize**

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

Analyze

We need the mass of 1 mol of carbon to compute the number of moles of carbon in the 500.0 g sample. From the periodic table, we see that the molar mass of carbon is 12.011 g/mol.

Solve

$$500.0 \text{ g C} \times \frac{1 \text{ mol}}{12.011 \text{ g}} = 41.63 \text{ mol C}$$

Think About It

Because carbon's molar mass is relatively low at 12 g/mol, 500 g of this substance contains a fairly substantial number of moles.

2.86. **Collect and Organize**

We are asked to convert a mass of gold in ounces to moles.

Analyze

We need the mass of 1 mol of gold to compute the number of moles of gold in the 2.00 oz sample. From the periodic table, we see that the molar mass of gold is 196.97 g/mol. We also must convert the mass of the gold in ounces to the mass in grams, using the conversion 1 oz = 28.35 g.

Solve

$$2.00 \text{ oz Au} \times \frac{28.35 \text{ g}}{1 \text{ oz}} \times \frac{1 \text{ mol}}{196.97 \text{ g}} = 0.288 \text{ mol Au}$$

Think About It

Because gold's molar mass is relatively high at 197 g/mol, a few ounces do not contain many moles of gold atoms.

2.87. **Collect and Organize**

Given a molar amount of calcium titanate, we are asked to determine the number of moles and mass 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.

2.88. **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 (or 1.7 mol for two significant digits)The mass of O^{2-} ions in the sample, therefore, is

$$1.7 \text{ mol } O^{2-} \times \frac{15.999 \text{ g}}{\text{mol}} = 27 \text{ g}$$

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.

2.89. **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 44 g/mol, and the molar mass of NO is 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.

2.90. **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 the molar mass 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 is 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 g/mol, and the formula mass for Na_2S is 77.9 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:

$$10\text{ g CrCl}_3 \times \frac{1\text{ mol}}{158\text{ g}} \times \frac{4\text{ ions}}{\text{CrCl}_3\text{ formula}} = 0.25\text{ mol ions in CrCl}_3$$
$$10\text{ g Na}_2\text{S} \times \frac{1\text{ mol}}{77.9\text{ g}} \times \frac{3\text{ ions}}{\text{Na}_2\text{S formula}} = 0.38\text{ mol ions in Na}_2\text{S}$$

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 molecular 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 = 58.5 g/mol, 0.34 mol ions in 10 g).

2.91. **Collect and Organize**

Given a mass of quartz, we are to determine the moles of SiO_2 present.

Analyze

To convert from mass to moles, we divide the mass given by the molar mass of SiO_2 [28.086 + 2(15.999) = 60.084 g/mol].

Solve

$$\frac{45.2\text{ g SiO}_2}{60.084\text{ g/mol}} = 0.752\text{ mol SiO}_2$$

Think About It

Because the initial mass is less than the molar mass, we would expect less than 1 mol of SiO_2 to be in the quartz sample.

2.92. **Collect and Organize**

Given a mass of halite, determine the moles of NaCl present.

Analyze

To convert from mass to moles, we divide the mass given by the molar mass of NaCl (22.990 + 35.453 = 58.443 g/mol).

Solve

$$\frac{6.82\text{ g NaCl}}{58.443\text{ g/mol}} = 0.117\text{ mol NaCl}$$

Think About It

The given mass of the halite crystal is about 1/10 that of the formula weight for NaCl, so we expect about 0.1 mol of NaCl to be present in the sample.

2.93. **Collect and Organize**

This exercise asks us to compute the moles of uranium and carbon (diamond) atoms in a 1 cm^3 block of each element and then to compare them.

Analyze

Starting with the 1 cm^3 block of each element, we can obtain the mass of the block by multiplying by the density of the element. Dividing that result by the molar mass of the element gives us the moles of atoms in that block. The element block with more moles of atoms must have more atoms. We can compute the actual number of atoms by multiplying the moles of atoms for each element by Avogadro's number.

Solve

$$1\text{ cm}^3\text{ C} \times \frac{3.514\text{ g}}{\text{cm}^3} \times \frac{1\text{ mol}}{12.011\text{ g}} \times \frac{6.022\times10^{23}\text{ C atoms}}{\text{mol}} = 1.762\times10^{23}\text{ atoms of C}$$

$$1\text{ cm}^3\text{ U} \times \frac{19.05\text{ g}}{\text{cm}^3} \times \frac{1\text{ mol}}{238.03\text{ g}} \times \frac{6.022\times10^{23}\text{ U atoms}}{\text{mol}} = 4.820\times10^{22}\text{ atoms of U}$$

Therefore, the 1 cm^3 block of diamond contains more atoms.

Think About It

We might expect that, because the block of uranium weighs so much more than the diamond block (more than five times as much), the uranium block would contain more atoms. However, we also have to take into account the very large molar mass of uranium. The result is that the diamond block has about 3.7 times more atoms in it than the same-sized block of uranium.

2.94. **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 density (Al = 1 mL/2.70 g, Sr = 1 mL/2.64 g). Since 1 mL = 1 cm^3, the length of a side of the cube for each element is the cube root of the volume.

Solve

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

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

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

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

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.

2.95. **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 call the *molecular ion*, M^+. This and other cations produces when the molecule fragments are separated base on their mass-to-charge (m/z) ratios and counted. 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.

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.

2.96. **Collect and Organize**

We are asked how isotopic abundances of HBr are reflected in its mass spectrum.

Analyze

Mass spectrometry plots the mass-to-charge ratio (*m*/*z*) versus relative intensity or relative abundance of the peaks. Because isotopes of the elements have different masses, we will be able to see isotopes that are present.

Solve

The isotopes that might be present in HBr would be ^1H (99.99% abundant), ^2H (0.01% abundant), ^{79}Br (50.69% abundant), and ^{81}Br (49.31% abundant). The abundance of ^2H would be too low to be detected by normal (not high resolution) mass spectrometry, so we anticipate that a sample of naturally occurring HBr would contain a mixture of $^1H^{79}Br$ and $^1H^{81}Br$, which would appear at *m*/*z* values of 80 and 82 amu, respectively, with the *m*/*z* peak for $^1H^{81}Br$ amu being slightly less intense.

Think About It

If ^2H were more abundant or our mass spectrometer were very sensitive, we would expect to see *m*/*z* ratios at 80 ($^1H^{79}Br$), 81 ($^2H^{79}Br$), 82 ($^1H^{81}Br$), and 83 ($^2H^{81}Br$) amu. Can you predict the relative intensities of those peaks?

2.97. **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 amu, the molecular mass of CO_2 is 44 amu and for C_3H_8 the molecular mass also is 44 amu. 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.

2.98. **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 the same m/z value because of their identical molecular masses, their fragmentation patterns will be very different.

Think About It

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

2.99. **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.

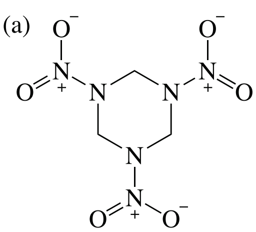
Solve

(a) $C_3H_6N_6O_6$: $(3 \times 12\ \text{amu}) + (6 \times 1\ \text{amu}) + (6 \times 14\ \text{amu}) + (6 \times 16\ \text{amu}) = 222\ \text{amu}$
(b) $C_4H_8N_8O_8$: $(4 \times 12\ \text{amu}) + (8 \times 1\ \text{amu}) + (8 \times 14\ \text{amu}) + (8 \times 16\ \text{amu}) = 296\ \text{amu}$
(c) $C_5H_8N_4O_{12}$: $(5 \times 12\ \text{amu}) + (8 \times 1\ \text{amu}) + (4 \times 14\ \text{amu}) + (12 \times 16\ \text{amu}) = 316\ \text{amu}$
(d) $C_{14}H_6N_6O_{12}$: $(14 \times 12\ \text{amu}) + (6 \times 1\ \text{amu}) + (6 \times 14\ \text{amu}) + (12 \times 16\ \text{amu}) = 450\ \text{amu}$

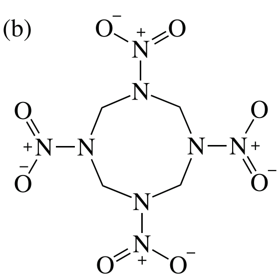
Think About It

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

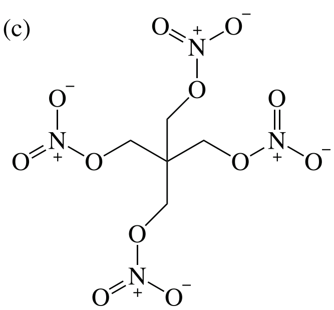
(a)



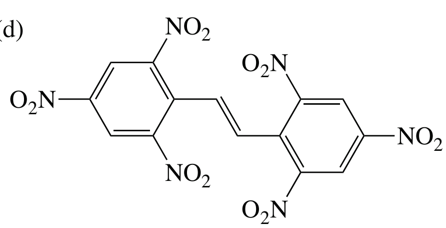
(b)



(c)



(d)



2.100. **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 × 12 amu) + (4 × 1 amu) = 16 amu
C_2H_6S: (2 × 12 amu) + (6 × 1 amu) + (1 × 32 amu) = 62 amu
$C_2H_2Cl_2$: (2 × 12 amu) + (2 × 1 amu) + (2 × 35.5 amu) = 97 amu

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?

2.101. **Collect and Organize**

From the mass spectrum for Cl_2 shown in Figure P2.101 and given the natural abundances of ^{35}Cl and ^{37}Cl, we are asked to explain why three peaks occur around *m*/*z* 72 amu and why the peak at 70 amu is taller than the one at 74 amu.

Analyze

The possible combinations of the chlorine isotopes in Cl_2 are $^{35}Cl_2$ (70 amu), $^{35}Cl^{37}Cl$ (72 amu), and $^{37}Cl_2$ (74 amu).

Solve

(a) The three peaks are due to the different combinations of isotopes in Cl_2: $^{35}Cl_2$ is at 70 amu, $^{35}Cl^{37}Cl$ is at 72 amu, and $^{37}Cl_2$ is at 74 amu.
(b) The peak at 70 amu is much more intense than the peak at 74 amu because ^{35}Cl is much more abundant than ^{37}Cl, so it is more likely that a molecule of Cl_2 will contain two ^{35}Cl isotopes than contain two ^{37}Cl isotopes.

Think About It

The mixed isotope can be made up in two ways: $^{35}Cl^{37}Cl$ and $^{37}Cl^{35}Cl$.

2.102. **Collect and Organize**

From the mass spectrum for Br_2 shown in Figure P2.102 and given the natural abundances of ^{79}Br and ^{81}Br, we are asked to explain why three peaks occur around *m*/*z* 160 amu and how we know that a third isotope of bromine, ^{80}Br, does not exist.

Analyze

The possible combinations of the bromine isotopes in Br_2 are $^{79}Br_2$ (158 amu), $^{79}Br^{81}Br$ (160 amu), and $^{81}Br_2$ (162 amu).

Solve

(a) The three peaks are due to the different combinations of isotopes in Br_2: $^{79}Br_2$ at 158 amu, $^{79}Br^{81}Br$ at 160 amu, and $^{81}Br_2$ at 162 amu.
(b) If ^{80}Br were present we would expect to see additional peaks in the mass spectrum: $^{79}Br^{80}Br$ at 159 amu and $^{80}Br^{81}Br$ at 161 amu, but we do not observe these in the mass spectrum.

Think About It

The peak for $^{80}Br_2$ would coincide with the $^{79}Br^{81}Br$ peak at 160 amu.

2.103. **Collect and Organize**

From the mass spectrum for H_2S shown in Figure P2.103 and given the natural abundances of four isotopes of sulfur, we are asked to explain the relative intensities as reflecting the abundance of the sulfur isotopes and the sequential loss of H from the molecule.

Analyze

The possible combinations of the sulfur isotopes in H_2S with their relative abundances are $H_2{}^{32}S$ (34 amu, 94.93%), $H_2{}^{33}S$ (35 amu, 0.76%), $H_2{}^{34}S$ (36 amu, 4.29%), and $H_2{}^{35}S$ (37 amu, 0.02%).

Solve

The biggest peak at 34 amu is the molecular ion of $H_2{}^{32}S$ (nearly 95% of all S atoms are ^{32}S). The peaks at 33 ns 32 amu are produced when electron bombardment of $H_2{}^{32}S$ molecules fragments them, forming ions $H^{32}S$ and ^{32}S atoms. The source of the small peak at 36 amu is probably the molecular ion of $H_2{}^{34}S$ (the second most abundant S isotope). The even smaller peak at 35 amu is probably a combination of mostly $H^{34}S$ and a very little $H_2{}^{33}S$ (this isotope of sulfur has a natural abundance of less than 1%). Molecular and fragment ions of the even less abundant ^{35}S isotope were probably not detected.

Think About It

Isotopic patterns such as these shown for H_2S are important in using mass spectrometry to identify compounds.

2.104. **Collect and Organize**

From the mass spectrum for AsH_3 shown in Figure P2.104 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 amu.

Solve

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

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.

2.105. **Collect and Organize**

From the mass spectrum of cocaine shown in Figure P2.105, 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 amu. 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.

2.106. **Collect and Organize**

From the mass spectrum of $(CH_3)_nSb$ shown in Figure P2.106 and given the natural abundances of ^{121}Sb and ^{123}Sb, we are to determine the number of CH_3 groups (the value of n) on this antimony molecule.

Analyze

The molecular ion peaks for this compound will give us a clue about the molar mass of the compound. These peaks are at 167 and 169 amu. The CH_3 unit has a mass of 15 amu.

Solve

We should be able to arrive at the value of n from either molecular ion peak. Using ^{121}Sb, which has a molecular ion mass at 167 amu, if we subtract the mass of the Sb isotope we get (167 – 121) = 46. Dividing that result by the mass of the CH_3 unit gives 46 ÷ 15 = 3. Therefore, the value of n is 3 and the formula for the molecule is $(CH_3)_3Sb$.

Think About It

You can see the successive loss of CH_3 units in the mass spectrum. The grouping of peaks occurs at around M^+–15, M^+–30, and M^+–45.

2.107. **Collect and Organize**

J. J. 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 were later called electrons.

Solve

(a) Today we call 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 2.2 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.

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.

2.108. **Collect and Organize**

First we are asked to determine the number of neutrons in each of four isotopes of strontium. Then, given the masses of the four isotopes of strontium and the abundances of two of them along with the average atomic mass, we are to calculate the natural abundances of the other two isotopes.

Analyze

(a) The number of neutrons in an isotope is the mass number minus the atomic number. For each isotope we subtract 38 from the mass number.
(b) To calculate the abundances of the two isotopes, we first realize that the total abundance for all four isotopes is 100%. That means that the abundance of ^{87}Sr + ^{88}Sr = 100 – (0.56 + 9.86) = 89.58%. We can then say that the abundance of ^{88}Sr = (89.58 – ^{87}Sr)%. We can substitute those expressions in to calculate the abundance of ^{87}Sr and then find the abundance of ^{88}Sr.

Solve

(a) Forty-six neutrons are in ^{84}Sr, 48 neutrons in ^{86}Sr, 49 neutrons in ^{87}Sr, and 50 neutrons in ^{88}Sr.
(b) Let the abundance of ^{87}Sr be x:

$$87.621 = (0.0056 \times 83.9134 \text{ amu}) + (0.0986 \times 85.9094 \text{ amu}) + (x \times 86.9089 \text{ amu}) + [(0.8958 - x) \times 87.9056 \text{ amu})$$
$$x = 0.0656$$

So the abundance of ^{87}Sr is 7.15%, and the abundance of ^{88}Sr is (89.58 – 6.56)% = 83.02%.

Think About It
The average mass is close to 88 amu, and that answer makes sense because the other isotopes are not in great abundance.

2.109. **Collect and Organize**
In this problem we are given the masses of the three isotopes of magnesium (^{24}Mg = 23.9850 amu, ^{25}Mg = 24.9858 amu, and ^{26}Mg = 25.9826 amu) and given that the abundance of ^{24}Mg is 78.99%. From that information and the average (weighted) atomic mass units of magnesium (24.3050 amu), 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

$$(0.7899 \times 23.9850) + 24.9858x + 25.9826y = 24.3050$$

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

$$(0.7899 \times 23.9850) + 24.9858(0.2101 - y) + 25.9826y = 24.3050$$

Solve

$$18.94575 + 5.249517 - 24.9858y + 25.9826y = 24.3050$$
$$0.9968y = 0.1097$$
$$y = 0.1101$$

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.

2.110. **Collect and Organize**
Using the periodic table in Figure P2.110, 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 number 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
Those elements are fluorine, sulfur, manganese, arsenic, and tantalum, respectively.

2.111. **Collect and Organize**
Given the diameter of 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 Avogadro's number, 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

$$1.00\ \text{g} \times \frac{1\ \text{mol}}{107.87\ \text{g}} \times \frac{6.022\times10^{23}\ \text{atoms}}{\text{mol}} \times \frac{1\ \text{nanoparticle}}{4.8\times10^{7}\ \text{atoms}} = 1.2\times10^{14}\ \text{nanoparticles}$$

Think About It

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

2.112. **Collect and Organize**

For CdS and CdSe quantum dots we are asked to calculate their formula masses, 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 molar masses of the elements Cd, Se, and S to calculate the formula masses of the CdSe and CdS quantum dots. Then we can use those formula masses with the molar masses of Se, Cd, and S to calculate the masses of Se and S in a 4.3×10^{-15} g sample of CdSe or CdS, respectively.

Solve

(a) Molar mass of CdS: 112.41 + 32.065 = 144.48 g/mol
Molar mass of CdSe: 112.41 + 78.96 = 191.37 g/mol

(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) $$4.3\times10^{-15}\ \text{g CdS} \times \frac{1\ \text{mol CdS}}{144.48\ \text{g}} \times \frac{1\ \text{mol Cd}}{1\ \text{mole CdS}} \times \frac{112.41\ \text{g Cd}}{1\ \text{mol Cd}} = 3.3\times10^{-15}\ \text{g Cd}$$

$$4.3\times10^{-15}\ \text{g CdS} \times \frac{1\ \text{mol CdS}}{144.48\ \text{g}} \times \frac{1\ \text{mol S}}{1\ \text{mole CdS}} \times \frac{32.065\ \text{g S}}{1\ \text{mol S}} = 9.5\times10^{-16}\ \text{g S}$$

Think About It

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

2.113. **Collect and Organize**

Given that the average molar mass of air is 28.8 g/mol and that each mole of air had 402.5×10^{-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

$$\frac{402.5\times10^{-6}\ \text{mol CO}_2}{\text{mol air}} \times \frac{\text{mol air}}{28.8\ \text{g air}} \times \frac{44.0\ \text{g CO}_2}{\text{mol CO}_2} \times \frac{1\ \mu\text{g CO}_2}{1\times10^{-6}\ \text{g CO}_2} = 615\ \mu\text{g CO}_2\ /\ \text{g air}$$

Think About It

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

2.114. **Collect and Organize**

From the mass spectrum shown in Figure P2.114, 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 mass of the compound by using values from the periodic table for C, H, and O.

Solve

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

Think About It

The molecular ion peak is not the most intense peak, but the peak of highest mass.

2.115. **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 Avogadro's number.

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.

2.116. **Collect and Organize**

Given that we know the abundances of three stable isotopes to the nearest 0.01% and we know the atomic masses to six significant figures, we are to determine to how many significant figures we will be able to compute the average atomic mass.

Analyze

We have to think back to the weakest-link principle in Chapter 1, and we have to make some assumptions here about the abundances. If we know the abundances to 0.01%, we would have four significant figures if all the elements are in equal abundance (33.33%). However, if one abundance is very low—say, 0.01%—then we would have one significant figure for that isotope.

Solve

Those two cases are in the extreme. For four significant figures for the weak link in the abundances, we would use four significant figures in multiplying the mass of each isotope by its abundance. Adding those would give us four significant figures in the average atomic mass. However, if we have only one significant figure for one of the abundances, the answer would have to be expressed in one significant figure (this is the other extreme).

Think About It
Remember that in multiplication or division we use the rule of the number of the least significant figures, and in addition or subtraction we use the rule of the least-certain digit in the measurement.