5.3 Molar Mass

Section Preview/ Specific Expectations

In this section, you will

- **explain** the relationship between the average atomic mass of an element and its molar mass
- **solve** problems involving number of moles, number of particles, and mass
- **calculate** the molar mass of a compound
- **communicate** your understanding of the following term: *molar mass*

In section 5.2, you explored the relationship between the number of atoms or particles and the number of moles in a sample. Now you are ready to relate the number of moles to the mass, in grams. Then you will be able to determine the number of atoms, molecules, or formula units in a sample by finding the mass of the sample.

Mass and the Mole

You would never express the mass of a lump of gold, like the one in Figure 5.11, in atomic mass units. You would express its mass in grams. How does the mole relate the number of atoms to measurable quantities of a substance? The definition of the mole pertains to relative atomic mass, as you learned in section 5.1. One atom of carbon-12 has a mass of exactly 12 u. Also, by definition, one mole of carbon-12 atoms (6.02×10^{23} carbon-12 atoms) has a mass of exactly 12 g.

The Avogadro constant is the factor that converts the relative mass of individual atoms or molecules, expressed in atomic mass units, to mole quantities, expressed in grams.

How can you use this relationship to relate mass and moles? The periodic table tells us the average mass of a single atom in atomic mass units (u). For example, zinc has an average atomic mass of 65.39 u. *One mole of an element has a mass expressed in grams numerically equivalent to the element's average atomic mass expressed in atomic mass units*. One mole of zinc atoms has a mass of 65.39 g. This relationship allows chemists to use a balance to count atoms. You can use the periodic table to determine the mass of one mole of an element.

Table 5.2 Average Atomic Mass and Molar Mass of Four Elements

Element	Average atomic mass (u)	Molar mass (g)
hydrogen, H	1.01	1.01
oxygen, O	16.00	16.00
sodium, Na	22.99	22.99
argon, Ar	39.95	39.95

What is Molar Mass?

The mass of one mole of any element, expressed in grams, is numerically equivalent to the average atomic mass of the element, expressed in atomic mass units. The mass of one mole of a substance is called its **molar mass** (symbol M). Molar mass is expressed in g/mol. For example, the average atomic mass of gold, as given in the periodic table, is 196.97 u. Thus the mass of one mole of gold atoms, gold's molar mass, is 196.97 g. Table 5.2 gives some additional examples of molar masses.

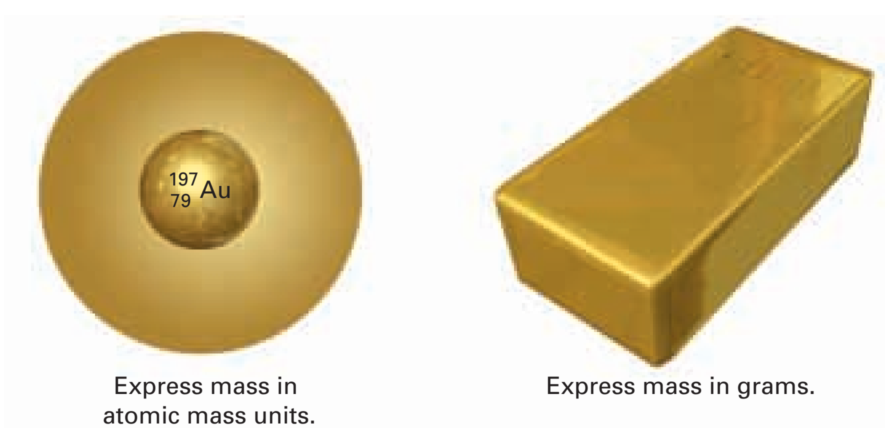

Figure 5.11 **The Avogadro constant is a factor that converts from atomic mass to molar mass.**

Finding the Molar Mass of Compounds

While you can find the molar mass of an element just by looking at the periodic table, you need to do some calculations to find the molar mass of a compound. For example, 1 mol of beryllium oxide, BeO, contains 1 mol of beryllium and 1 mol of oxygen. To find the molar mass of BeO, add the mass of each element that it contains.

$$\begin{aligned} M_{\text{BeO}} &= 9.01 \text{ g/mol} + 16.00 \text{ g/mol} \\ &= 25.01 \text{ g/mol} \end{aligned}$$

Examine the following Sample Problem to learn how to determine the molar mass of a compound. Following Investigation 5-A on the next page, there are some Practice Problems for you to try.

CHEM FACT

The National Institute of Standards and Technology (NIST) and most other standardization bodies use *M* to represent molar mass. You may see other symbols, such as *mm*, used to represent molar mass.

Sample Problem

Molar Mass of a Compound

Problem

What is the mass of one mole of calcium phosphate, $Ca_3(PO_4)_2$?

What Is Required?

You need to find the molar mass of calcium phosphate.

What Is Given?

You know the formula of calcium phosphate. You also know, from the periodic table, the average atomic mass of each atom that makes up calcium phosphate.

Plan Your Strategy

Find the total mass of each element to determine the molar mass of calcium phosphate. Find the mass of 3 mol of calcium, the mass of 2 mol of phosphorus, and the mass of 8 mol of oxygen. Then add these masses together.

Act on Your Strategy

$M_{\text{Ca}} \times 3 = (40.08 \text{ g/mol}) \times 3 = 120.24 \text{ g/mol}$

$M_{\text{P}} \times 2 = (30.97 \text{ g/mol}) \times 2 = 61.94 \text{ g/mol}$

$M_{\text{O}} \times 8 = (16.00 \text{ g/mol}) \times 8 = 128.00 \text{ g/mol}$

$$\begin{aligned} M_{\text{Ca}_3(\text{PO}_4)_2} &= 120.24 \text{ g/mol} + 61.94 \text{ g/mol} + 128.00 \text{ g/mol} \\ &= 310.18 \text{ g/mol} \end{aligned}$$

Therefore the molar mass of calcium phosphate is 310.18 g/mol.

Check Your Solution

Using round numbers for a quick check, you get

$$(40 \times 3) + (30 \times 2) + (15 \times 8) = 300$$

This estimate is close to the answer of 310.18 g/mol.

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PROBLEM TIP

Once you are used to calculating molar masses, you will want to do the four calculations at left all at once. Try solving the Sample Problem using only one line of calculations.

SKILL FOCUS

Modelling concepts

Analyzing and interpreting

Communicating results

Investigation 5-A

Modelling Mole and Mass Relationships

Chemists use the mole to group large numbers of atoms and molecules into manageable, macroscopic quantities. In this way, they can tell how many atoms or molecules are in a given sample, even though the particles are too small to see. In this investigation, you will explore how to apply this idea to everyday objects such as grains of rice and nuts and bolts.

Question

How can you use what you know about the mole and molar mass to count large numbers of tiny objects using mass, and to relate numbers of objects you cannot see based on their masses?

Part 1 Counting Grains of Rice

Materials

electronic balance

40 mL dry rice

50 mL beaker

Procedure

1. Try to measure the mass of a grain of rice. Does the balance register this mass?
2. Count out 20 grains of rice. Measure and record their mass.
3. Find the mass of the empty beaker. Add the rice to the beaker. Find the mass of the beaker and the rice. Determine the mass of the rice.
4. Calculate the number of grains of rice in the 40 mL sample. Report your answer to the number of significant digits that reflects the precision of your calculation.

Part 2 Counting Objects Based on Their Relative Masses

Materials

electronic balance

10 small metal nuts (to represent the fictitious element *nutium*)

10 washers (to represent the fictitious element *washerium*)

2 opaque film canisters with lids

Procedure

1. Measure the mass of 10 nuts (*nutium* atoms). Then measure the mass of 10 washers (*washerium* atoms).
2. Calculate the average mass of a single "atom" of *nutium* and *washerium*.
3. Determine the mass ratio of *nutium* to *washerium*.
4. Obtain, from your teacher, a sealed film canister containing an unknown number of *nutium* atoms. Your teacher will tell you the mass of the empty film canister and lid.
5. Find the mass of the unknown number of *nutium* atoms.
6. You know that you need an equal number of *washerium* atoms to react with the unknown number of *nutium* atoms. What mass of *washerium* atoms do you need?

Analysis

1. Was it possible to get an accurate mass for an individual grain of rice? How did you solve this problem?
2. How did you avoid having to count every single grain of rice in order to determine how many there were in the sample?
3. Using your data, how many grains of rice would be in 6.5×10^3 g of rice?
4. A mole of helium atoms weighs 4.00 g.
   (a) How many atoms are in 23.8 g of helium?
   (b) What known relationship did you use to find your answer to part (a)?
   (c) What analogous relationship did you set up in order to calculate the number of grains of rice based on the mass of the rice?
5. You know the relative mass of nuts and washers. Suppose that you are given some washers in a sealed container. You know that you have the same number of nuts in another sealed container.
   (a) Can you determine how many washers are in the container without opening either container? Why or why not?
   (b) What, if any, additional information do you need?
6. The molar mass of carbon is 12.0 g. The molar mass of molecular oxygen is 32.0 g. Equal numbers of carbon atoms and oxygen molecules react to form carbon dioxide.
   (a) If you have 5.8 g of carbon, what mass of oxygen will react?
   (b) How does part (a) relate to step 6 in the Procedure for Part 2?

Conclusion

7. How do chemists use the mole and molar masses to count numbers and relative numbers of atoms and molecules? Relate your answer to the techniques you used to count rice, nuts, and washers.

Applications

8. Think about your answer to Analysis question 5(a). Did you need to use the Avogadro constant in your calculation? Explain why or why not.
9. Chemists rarely use the Avogadro constant directly in their calculations. What relationship do they use to avoid working with such a large number?

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Practice Problems

23. State the molar mass of each element.

(a) xenon, Xe

(b) osmium, Os

(c) barium, Ba

(d) tellurium, Te

24. Find the molar mass of each compound.

(a) ammonia, NH_3

(b) glucose, $C_6H_{12}O_6$

(c) potassium dichromate, $K_2Cr_2O_7$

(d) iron(III) sulfate, $Fe_2(SO_4)_3$

25. Strontium may be found in nature as celestite, $SrSO_4$. Find the molar mass of celestite.

26. What is the molar mass of the ion $[Cu(NH_3)_4]^{2+}$?

Counting Particles Using Mass

Using the mole concept and the periodic table, you can determine the mass of one mole of a compound. You know, however, that one mole represents 6.02×10^{23} particles. Therefore you can use a balance to count atoms, molecules, or formula units!

For example, consider carbon dioxide, CO_2. One mole of carbon dioxide has a mass of 44.0 g and contains 6.02×10^{23} molecules. You can set up the following relationship:

$$6.02 \times 10^{23} \text{ molecules of } CO_2 \rightarrow 1 \text{ mol of } CO_2 \rightarrow 44.0 \text{ g of } CO_2$$

How can you use this relationship to find the number of molecules and the number of moles in 22.0 g of carbon dioxide?

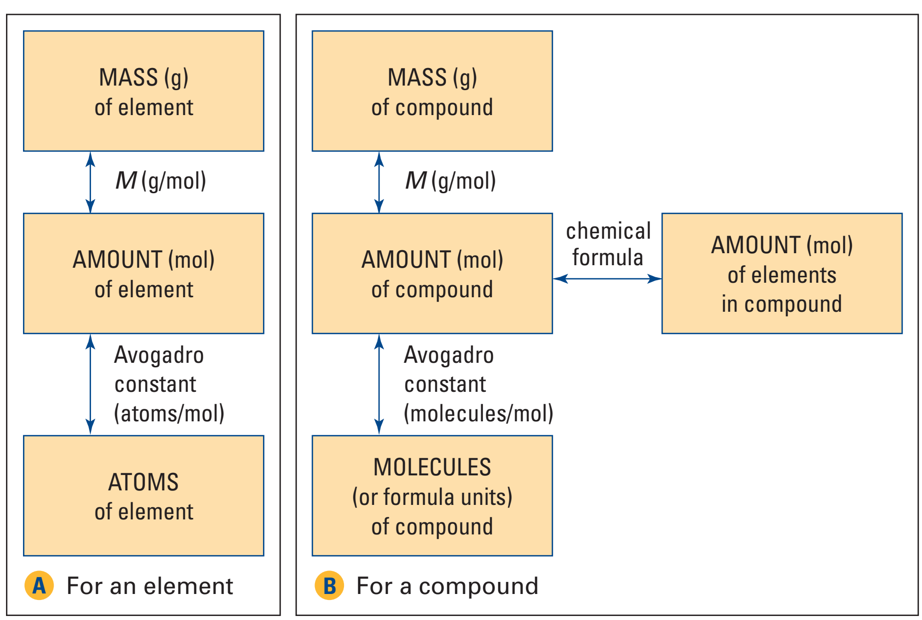

Figure 5.12 The molar mass relates the amount of an element or a compound, in moles, to its mass. Similarly, the Avogadro constant relates the number of particles to the molar amount.

Converting from Moles to Mass

Suppose that you want to carry out a reaction involving ammonium sulfate and calcium chloride. The first step is to obtain one mole of each chemical. How do you decide how much of each chemical you need? You convert the molar amount to mass. Then you use a balance to determine the mass of the proper amount of each chemical.

The following equation can be used to solve problems involving mass, molar mass, and number of moles:

$$\text{Mass} = \text{Number of moles} \times \text{Molar mass}$$
$$m = n \times M$$

$$\boxed{\text{moles (mol)}} \xrightarrow[\text{molar mass (g/mol).}]{\text{Multiply by}} \boxed{\text{mass (g)}}$$

CHECKPOINT

Write the chemical formulas for calcium chloride and ammonium sulfate. Predict what kind of reaction will occur between them. Write a balanced chemical equation to show the reaction. Ionic compounds containing the ammonium ion are soluble. Ammonium sulfate is soluble, but barium chloride is not.

Sample Problem

Moles to Mass

Problem

A flask contains 0.750 mol of carbon dioxide gas, CO_2. What mass of carbon dioxide gas is in this sample?

What Is Required?

You need to find the mass of carbon dioxide.

What Is Given?

The sample contains 0.750 mol. You can determine the molar mass of carbon dioxide from the periodic table.

Plan Your Strategy

In order to convert moles to grams, you need to determine the molar mass of carbon dioxide from the periodic table.

Multiply the molar mass of carbon dioxide by the number of moles of carbon dioxide to determine the mass.

$$m = n \times M$$

Act on Your Strategy

$$M_{CO_2} = 2 \times (16.00 \text{ g/mol}) + 12.01 \text{ g/mol}$$
$$= 44.01 \text{ g/mol}$$

$$m = (0.750 \cancel{\text{mol}}) \times (44.01 \text{ g/}\cancel{\text{mol}})$$
$$= 33.0 \text{ g}$$

The mass of 0.75 mol of carbon dioxide is 33.0 g.

Check Your Solution

1 mol of carbon dioxide has a mass of 44 g. You need to determine the mass of 0.75 mol, or 75% of a mole. 33 g is equal to 75% of 44 g.

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Practice Problems

27. Calculate the mass of each molar quantity.

(a) 3.90 mol of carbon, C

(b) 2.50 mol of ozone, O_3

(c) 1.75×10^{7} mol of propanol, C_3H_8O

(d) 1.45×10^{-5} mol of ammonium dichromate, $(NH_4)_2Cr_2O_7$

28. For each group, which sample has the largest mass?

(a) 5.00 mol of C, 1.50 mol of Cl_2, 0.50 mol of $C_6H_{12}O_6$

(b) 7.31 mol of O_2, 5.64 mol of CH_3OH, 12.1 mol of H_2O

29. A litre, 1000 mL, of water contains 55.6 mol. What is the mass of a litre of water?

30. To carry out a particular reaction, a chemical engineer needs 255 mol of styrene, C_8H_8. How many kilograms of styrene does the engineer need?

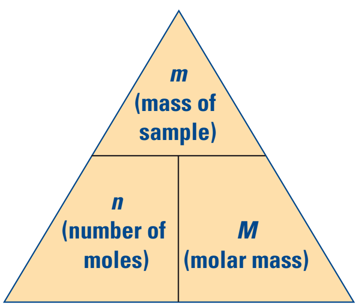

Figure 5.13 Use this triangle for problems involving number of moles, mass of sample, and molar mass. For what other scientific relationships might you use a triangle like this?

You might find the triangle shown in Figure 5.13 useful for problems involving number of moles, number of particles, and molar mass. To use it, cover the quantity that you need to find. The required operation—multiplication or division—will be obvious from the position of the remaining variables. For example, if you want to find the mass of a sample, cover the *m* in the triangle. You can now see that

$$\text{Mass} = \text{Number of moles} \times \text{Molar mass.}$$

Be sure to check that your units cancel.

Converting from Mass to Moles

In the previous Sample Problem, you saw how to convert moles to mass. Often, however, chemists know the mass of a substance but are more interested in knowing the number of moles. Suppose that a reaction produces 223 g of iron and 204 g of aluminum oxide. The masses of the substances do not tell you very much about the reaction. You know, however, that 223 g of iron is 4 mol of iron. You also know that 204 g of aluminum oxide is 2 mol of aluminum oxide. You may conclude that the reaction produces twice as many moles of iron as it does moles of aluminum oxide. You can perform the reaction many times to test your conclusion. If your conclusion is correct, the mole relationship between the products will hold. To calculate the number of moles in a sample, find out how many times the molar mass goes into the mass of the sample.

$$\text{moles (mol)} \xleftarrow[\text{molar mass (g/mol).}]{\text{Divide by}} \text{mass (g)}$$

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

$$n = \frac{m}{M}$$

The following Sample Problem explains how to convert from the mass of a sample to the number of moles it contains.

Sample Problem

Mass to Moles

Problem

How many moles of acetic acid, CH_3COOH, are in a 23.6 g sample?

What Is Required?

You need to find the number of moles in 23.6 g of acetic acid.

What Is Given?

You are given the mass of the sample.

Plan Your Strategy

To obtain the number of moles of acetic acid, divide the mass of acetic acid by its molar mass.

Act on Your Strategy

The molar mass of CH_3COOH is
$(12.01 \times 2) + (16.00 \times 2) + (1.01 \times 4) = 60.06$ g.

$$n = \frac{m}{M_{CH_3COOH}}$$

$$n \text{ mol } CH_3COOH = \frac{23.6 \cancel{\text{g}}}{60.06 \cancel{\text{g}}/\text{mol}}$$

$$= 0.393 \text{ mol}$$

Therefore there are 0.393 mol of acetic acid in 23.6 g of acetic acid.

Check Your Solution

Work backwards. There are 60.06 g in each mol of acetic acid. So in 0.393 mol of acetic acid, you have $0.393 \cancel{\text{mol}} \times 60.06 \text{ g}/\cancel{\text{mol}} = 23.6$ g of acetic acid. This value matches the question.

Practice Problems

31. Calculate the number of moles in each sample.

(a) 103 g of Mo

(b) 1.32×10^4 g of Pd

(c) 0.736 kg of Cr

(d) 56.3 mg of Ge

32. How many moles of compound are in each sample?

(a) 39.2 g of silicon dioxide, SiO_2

(b) 7.34 g of nitrous acid, HNO_2

(c) 1.55×10^5 kg of carbon tetrafluoride, CF_4

(d) 8.11×10^{-3} mg of 1-iodo-2,3-dimethylbenzene C_8H_9I

33. Sodium chloride, NaCl, can be used to melt snow. How many moles of sodium chloride are in a 10 kg bag?

34. Octane, C_8H_{18}, is a principal ingredient of gasoline. Calculate the number of moles in a 20.0 kg sample of octane.

Chemistry Bulletin

Science | Technology | Society | Environment

Chemical Amounts in Vitamin Supplements

Vitamins and minerals (micronutrients) help to regulate your metabolism. They are the building blocks of blood and bone, and they maintain muscles and nerves. In Canada, a standard called *Recommended Nutrient Intake* (*RNI*) outlines the amounts of micronutrients that people should ingest each day. Eating a balanced diet is the best way to achieve your RNI. Sometimes, however, you may need to take multivitamin supplements when you are unable to attain your RNI through diet alone.

The label on a bottle of supplements lists all the vitamins and minerals the supplements contain. It also lists the form and source of each vitamin and mineral, and the amount of each. The form of a mineral is especially important to know because it affects the quantity your body can use. For example, a supplement may claim to contain 650 mg of calcium carbonate, $CaCO_3$, per tablet. This does not mean that there is 650 mg of calcium. The amount of actual calcium, or *elemental calcium*, in calcium carbonate is only 260 mg. Calcium carbonate has more elemental calcium than the same amount of calcium gluconate, which only has 58 mg for every 650 mg of the compound. Calcium gluconate may be easier for your body to absorb, however.

Quality Control

Multivitamin manufacturers employ chemists, or analysts, to ensure that the products they make have the right balance of micronutrients. Manufacturers have departments devoted to *quality control* (*QC*). QC chemists analyze all the raw materials in the supplements, using standardized tests. Most manufacturers use tests approved by a "standardization body," such as the US Pharmacopoeia. Such standardization bodies have developed testing guidelines to help manufacturers ensure that their products contain what the labels claim, within strict limits.

To test for quality, QC chemists prepare samples of the raw materials from which they will make the supplements. They label the samples according to the "lot" of materials from which the samples were taken. They powder and weigh the samples. Then they extract the vitamins. At the same time, they prepare standard solutions containing a known amount of each vitamin.

Next the chemists compare the samples to the standards by subjecting both to the same tests. One test that is used is *high-performance liquid chromatography* (*HPLC*). HPLC produces a spectrum, or "fingerprint," that identifies each compound. Analysts compare the spectrum that is produced by the samples to the spectrum that is produced by the standard.

Analysts test tablets and capsules for dissolution and disintegration properties. The analyst may use solutions that simulate the contents of the human stomach or intestines for these tests. Only when the analysts are sure that the tablets pass all the necessary requirements are the tablets shipped to retail stores.

Making Connections

1. Why might consuming more of the daily RNI of a vitamin or mineral be harmful?
2. The daily RNI of calcium for adolescent females is 700 to 1100 mg. A supplement tablet contains 950 mg of calcium citrate. Each gram of calcium citrate contains 5.26×10^{-3} mol calcium. How many tablets would a 16-year-old female have to take to meet her daily RNI?

An analyst tests whether tablets will sufficiently dissolve within a given time limit.

Converting Between Moles, Mass, and Number of Particles

You can use what you now know about the mole to carry out calculations involving molar mass and the Avogadro constant. One mole of any compound or element contains 6.02×10^{23} particles. The compound or element has a mass, in grams, that is determined from the periodic table.

Now that you have learned how the number of particles, number of moles, and mass of a substance are related, you can convert from one value to another. Usually chemists convert from moles to mass and from mass to moles. Mass is a property that can be measured easily. The following graphic shows the factors used to convert between particles, moles, and mass. Moles are a convenient way to communicate the amount of a substance.

$$\text{number of particles} \xleftrightarrow[\text{constant}]{\text{Avogadro}} \text{moles (mol)} \xleftrightarrow[\text{(g/mol)}]{\text{molar mass}} \text{mass (g)}$$

For example, suppose that you need 2.3 mol of potassium chloride to carry out a reaction. You need to convert the molar amount to mass so that you can measure the correct amount with a balance.

To be certain you understand the relationship among particles, moles, and mass, examine the following Sample Problem.

Sample Problem

Particles to Mass

Problem

What is the mass of 5.67×10^{24} molecules of cobalt(II) chloride, $CoCl_2$?

What Is Required?

You need to find the mass of 5.67×10^{24} molecules of cobalt(II) chloride.

What Is Given?

You are given the number of molecules.

Plan Your Strategy

Convert the number of molecules into moles by dividing by the Avogadro constant. Then convert the number of moles into grams by multiplying by the molar mass of cobalt(II) chloride.

Act on Your Strategy

$$\frac{\text{Number of molecules } CoCl_2}{\text{Number of molecules } CoCl_2/\text{mol } CoCl_2} \times \text{mass } CoCl_2/\text{mol } CoCl_2$$

$$= \frac{5.67 \times 10^{24} \cancel{\text{ molecules } CoCl_2}}{6.02 \times 10^{23} \cancel{\text{ molecules } CoCl_2}/\cancel{\text{mol } CoCl_2}} \times 129.84 \text{ g } CoCl_2/\cancel{\text{mol } CoCl_2}$$

$$= 1.22 \times 10^{3} \text{ g } CoCl_2$$

Continued ...

Math LINK

Average atomic mass values in some periodic tables can have five or more significant digits. How do you know how many significant digits to use? When you enter values, such as average atomic mass, into your calculator, be sure that you use at least one more significant digit than is required in your answer.

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Check Your Solution

5.67×10^{24} molecules is roughly 10 times the Avogadro constant. This means that you have about 10 mol of cobalt(II) chloride. The molar mass of cobalt(II) chloride is about 130 g, and 10 times 130 g is 1300 g.

Practice Problems

35. Determine the mass of each sample.

(a) 6.02×10^{24} formula units of $ZnCl_2$

(b) 7.38×10^{21} formula units of $Pb_3(PO_4)_2$

(c) 9.11×10^{23} molecules of $C_{15}H_{21}N_3O_{15}$

(d) 1.20×10^{29} molecules of N_2O_5

36. What is the mass of lithium in 254 formula units of lithium chloride, LiCl?

37. Express the mass of a single atom of titanium, Ti, in grams.

38. Vitamin B_2, $C_{17}H_{20}N_4O_6$, is also called riboflavin. What is the mass, in grams, of a single molecule of riboflavin?

What if you wanted to compare amounts of substances, and you only knew their masses? You would probably convert their masses to moles. The Avogadro constant relates the molar amount to the number of particles. Examine the next Sample Problem to learn how to convert mass to number of particles.

Sample Problem

Mass to Particles

Problem

Chlorine gas, Cl_2, can react with iodine, I_2, to form iodine chloride, ICl. How many molecules of iodine chloride are contained in a 2.74×10^{-1} g sample?

What Is Required?

You need to find the number of molecules in 2.74×10^{-1} g of iodine.

What Is Given?

You are given the mass of the sample.

Continued ...

Continued ...
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Plan Your Strategy

First convert the mass to moles, using the molar mass of iodine. Multiplying the number of moles by the Avogadro constant will yield the number of molecules.

Act on Your Strategy

The molar mass of ICl is 162.36 g.
Dividing the given mass of ICl by the molar mass gives

$$n = \frac{2.74 \times 10^{-1}\ \cancel{g}}{162.36\ \cancel{g}/\text{mol}}$$
$$= 1.69 \times 10^{-3}\ \text{mol}$$

Now multiply the number of moles by the Avogadro constant. This gives the number of molecules in the sample.

$$(1.69 \times 10^{-3}\ \cancel{\text{mol}}) \times \frac{(6.02 \times 10^{23}\ \text{molecules})}{1\ \cancel{\text{mol}}} = 1.01 \times 10^{21}\ \text{molecules}$$

Therefore there are 1.01×10^{21} molecules in 2.74×10^{-1} g of iodine chloride.

Check Your Solution

Work backwards. Each mole of iodine chloride has a mass of 162.36 g/mol. Therefore 1.01×10^{21} molecules of iodine chloride have a mass of:

$$1.01 \times 10^{21}\ \cancel{\text{molecules}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}} \times \frac{162.36\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$= 2.72 \times 10^{-1}\ \text{g}$$

The answer is close to the value given in the question. Your answer is reasonable.

Practice Problems

39. Determine the number of molecules or formula units in each sample.

(a) 10.0 g of water, H_2O

(b) 52.4 g of methanol, CH_3OH

(c) 23.5 g of disulfur dichloride, S_2Cl_2

(d) 0.337 g of lead(II) phosphate, $Pb_3(PO_4)_2$

40. How many atoms of hydrogen are in 5.3×10^4 molecules of sodium glutamate, $NaC_5H_8NO_4$?

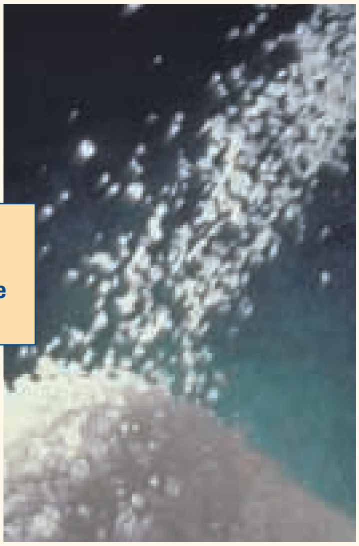
41. How many molecules are in a 64.3 mg sample of tetraphosphorus decoxide, P_4O_{10}?

42. **(a)** How many formula units are in a 4.35×10^{-2} g sample of potassium chlorate, $KClO_3$?

(b) How many ions (chlorate and potassium) are in this sample?

Concept Organizer Relationships Among Particles, Moles, and Mass

$$\begin{tikzcd}[column sep=huge]
\text{Number of particles} \arrow[r, shift left, "\text{divide by the Avogadro constant}"] & \text{Moles} \arrow[l, shift left, "\text{multiply by the Avogadro constant}"] \arrow[r, shift left, "\text{multiply by molar mass}"] & \text{Mass of substance} \arrow[l, shift left, "\text{divide by molar mass}"]
\end{tikzcd}$$



Section Wrap-up

In this chapter, you have learned about the relationships among the number of particles in a substance, the amount of a substance in moles, and the mass of a substance. Given the mass of any substance, you can now determine how many moles and particles make it up. In the next chapter, you will explore the mole concept further. You will learn how the mass proportions of elements in compounds relate to their formulas.

Section Review

Unit Investigation Prep

Before you design your experiment to determine the composition of a mixture, be sure you understand the relationship between moles and mass.

1. C Draw a diagram that shows the relationship between the atomic mass and molar mass of an element and the Avogadro constant.
2. I Consider a 78.6 g sample of ammonia, NH_3.
   (a) How many moles of ammonia are in the sample?
   (b) How many molecules of ammonia are in the sample?
3. I Use your understanding of the mole to answer the following questions.
   (a) What is the average mass, in grams, of a single atom of silicon, Si?
   (b) What is the mass, in atomic mass units, of a mole of silicon atoms?
4. I Consider a 0.789 mol sample of sodium chloride, NaCl.
   (a) What is the mass of the sample?
   (b) How many formula units of sodium chloride are in the sample?
   (c) How many ions are in the sample?
5. I A 5.00 carat diamond has a mass of 1.00 g. How many carbon atoms are in a 5.00 carat diamond?
6. I A bottle of mineral supplement tablets contains 100 tablets and 200 mg of copper. The copper is found in the form of cupric oxide. What mass of cupric oxide is contained in each tablet?