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10 CHEMICAL QUANTITIES

SECTION 10.1 THE MOLE: A MEASUREMENT OF MATTER (pages 287–296)

This section defines the mole and explains how the mole is used to measure matter. It also teaches you how to calculate the mass of a mole of any substance.

▶ Measuring Matter (pages 287–289)

1. What do the questions "how much?" and "how many?" have in common?

   They are questions about the amount of a substantce and are similar to questions

   scientists ask.

2. List two or three ways to measure matter.

   count the matter, measure the mass or weight, measure the volume

▶ What Is a Mole? (pages 290–293)

3. Circle the letter of the term that is an SI unit for measuring the amount of a substance.

   a. dozen   **b.** ounce   **c.** pair   **(d.)** mole

4. What is Avogadro's number?

   6.02×10^{23} representative particles of a substance

5. Circle the letter of the term that is NOT a representative particle of a substance.

   a. molecule   **b.** atom   **(c.)** grain   **d.** formula unit

6. List the representative particle for each of the following types of substances.

   a. molecular compounds   _____molecule_____

   b. ionic compounds   _____formula unit_____

   c. elements   _____atom_____

7. Is the following sentence true or false? To determine the number of

   representative particles in a compound, you count the molecules by viewing

   them under a microscope. _____false_____

8. How can you determine the number of atoms in a mole of a molecular compound?

   Use the chemical formula to find the number of atoms in one molecule and multiply

   this number by Avogadro's number, the number of particles in one mole.

CHAPTER 10, Chemical Quantities *(continued)*

9. Complete the table about representative particles and moles.

Representative Particles and Moles			
	Representative Particle	**Chemical Formula**	**Representative Particles in 1.00 mol**
Atomic oxygen	Atom	O	6.02×10^{23}
Oxygen gas	Molecule	O_2	6.02×10^{23}
Sodium ion	Ion	Na^+	6.02×10^{23}
Sodium chloride	Formula unit	NaCl	6.02×10^{23}

▶ The Mass of a Mole of an Element (pages 293–294)

10. What is the atomic mass of an element?

The atomic mass of an element is the mass of a single atom in atomic mass units.

11. Circle the letter of the phrase that completes this sentence correctly.
The atomic masses of all elements

a. are the same.

b. are based on the mass of the carbon isotope C-12.

c. are based on the mass of a hydrogen atom.

▶ The Mass of a Mole of a Compound (pages 295–296)

12. How do you determine the mass of a mole of a compound?

The mass of a mole of a compound is determined by adding the atomic masses of

the atoms making up the molecule.

13. Complete the labels on the diagram below.

SO_3 = 1 S atom + 3 O atoms

80.1 amu = 32.1 amu + 48.0 amu

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14. What is the molar mass of a compound?

It is the mass of 1 mol of that compound.

15. Is the following sentence true or false? Molar masses can be calculated directly from atomic masses expressed in grams. _____true_____

SECTION 10.2 MOLE–MASS AND MOLE–VOLUME RELATIONSHIPS (pages 297–303)

This section explains how to use molar mass and molar volume to convert among measurements of mass, volume, and number of particles.

▶ The Mole–Mass Relationship (pages 297–299)

1. What is the molar mass of a substance?

It is the mass (in grams) of one mole of the substance.

2. What is the molar mass of KI (potassium iodide)?

39.1 g K + 126.9 g I = 166.0 g KI

▶ The Mole–Volume Relationship (pages 300–302)

3. Is the following sentence true or false? The volumes of one mole of different solid and liquid substances are the same. _____false_____

4. Circle the letter of each term that can complete this sentence correctly. The volume of a gas varies with a change in

(a.) temperature.   **(c.)** pressure.

(b.) the size of the container.   **d.** the amount of light in the container.

5. Circle the letter of the temperature that is defined as standard temperature.

a. 0 K   **(c.)** 0°C

b. 100 K   **d.** 100°C

6. Is the following sentence true or false? Standard pressure is 101.3 kPa or 1 atmosphere (atm). _____true_____

7. What is the molar volume of a gas at standard temperature and pressure (STP)? _____22.4 L_____

8. What units do you normally use to describe the density of a gas?

_____grams per liter (g/L)_____

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CHAPTER 10, Chemical Quantities *(continued)*

9. What is Avogadro's hypothesis?

Avogadro's hypothesis says that equal volumes of gases at the same temperature

and pressure contain equal numbers of particles.

10. Look at Figure 10.9 on page 300 to help you answer this question. Why is Avogadro's hypothesis reasonable?

As long as the gas particles are not tightly packed, there is a great deal of empty

space between them. A container can easily accommodate the same number of

relatively large or relatively small gas particles.

11. How many gas particles occupy a volume of 22.4 L at standard temperature

and pressure? 6.02×10^{23} particles

▶ The Mole Road Map (page 303)

12. The figure below shows how to convert from one unit to another unit. Write the missing conversion factors below.

.

a. $\dfrac{1.00 \text{ mol}}{22.4 \text{ L}}$
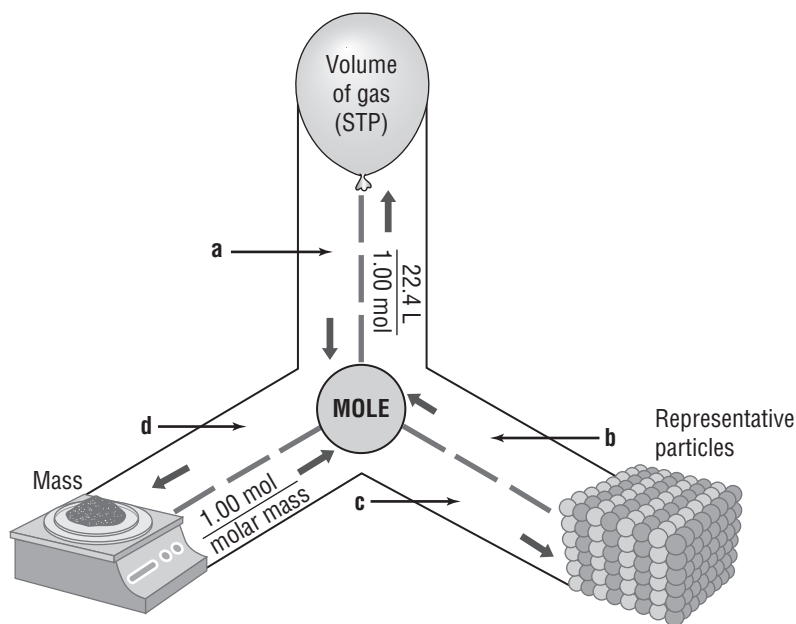
c. $\dfrac{6.02 \times 10^{23} \text{ particles}}{1.00 \text{ mol}}$

b. $\dfrac{1.00 \text{ mol}}{6.02 \times 10^{23} \text{ particles}}$

d. $\dfrac{\text{molar mass}}{1.00 \text{ mol}}$

SECTION 10.3 PERCENT COMPOSITION AND CHEMICAL FORMULAS (pages 305–312)

This section explains how to calculate percent composition from chemical formulas or experimental data, and how to derive empirical and molecular formulas.

▶ Percent Composition of a Compound (pages 305–308)

1. How do you express relative amounts of each element in a compound?

Relative amounts are expressed by the percent composition or the percent by mass.

2. Circle the letter of the phrase that completes this sentence correctly. The number of percent values in the percent composition of a compound is

   a. half as many as there are different elements in the compound.

   (b.) as many as there are different elements in the compound.
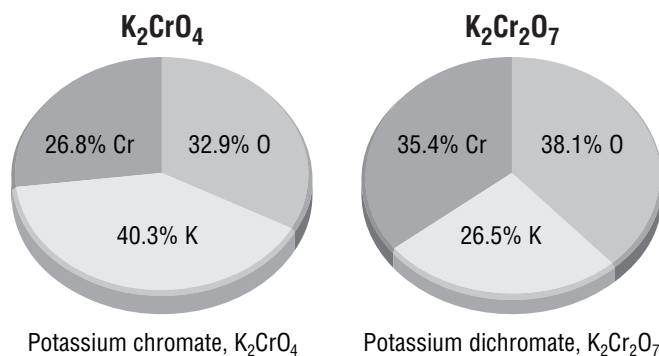
   c. twice as many as there are different elements in the compound.

3. What is the formula for the percent by mass of an element in a compound?

$$\% \text{ mass of element} = \frac{\text{grams of element}}{\text{grams of compound}} \times 100\%$$

4. In the diagram below, which compound has a greater percent composition of chromium? _____potassium dichromate_____

How much greater is this percent? _____8.6%_____

K_2CrO_4

26.8% Cr   32.9% O

40.3% K

Potassium chromate, K_2CrO_4

$K_2Cr_2O_7$

35.4% Cr   38.1% O

26.5% K

Potassium dichromate, $K_2Cr_2O_7$

5. To calculate the percent composition of a known compound, start with the chemical formula of the compound and calculate the _____molar mass_____ , which gives the mass of one mole of the compound.

6. Is the following sentence true or false? You can use percent composition to calculate the number of grams of an element in a given amount of a compound. _____true_____

CHAPTER 10, Chemical Quantities *(continued)*

7. How do you calculate the grams of an element in a specific amount of a compound?
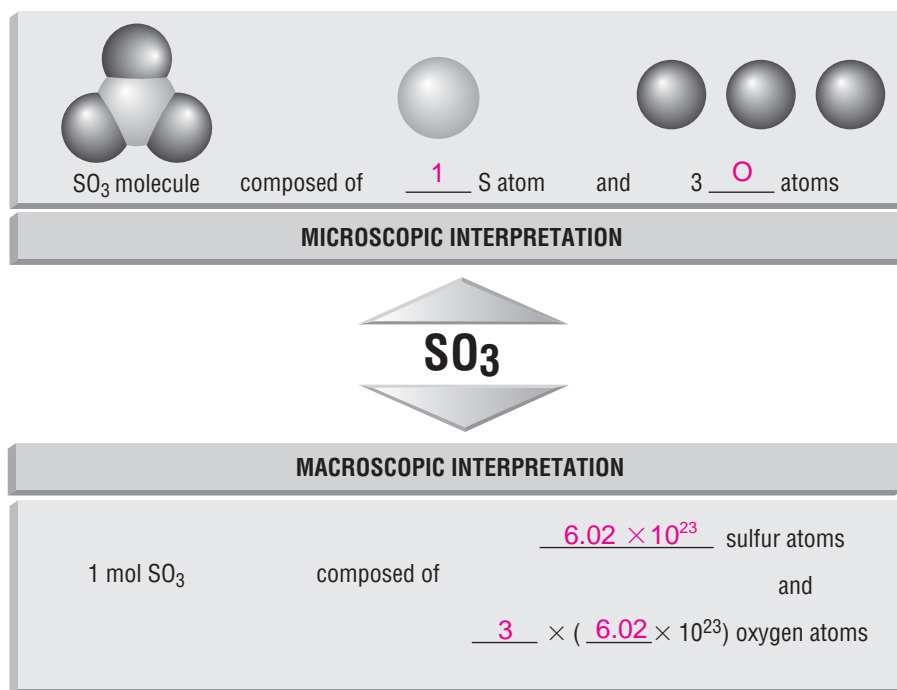
Multiply the mass of the compound by a conversion factor that is based on the

percent composition.

▶ Empirical Formulas (pages 309–310)

8. An empirical formula of a compound gives the _____lowest_____

whole-number ratio of the atoms of the elements in a compound.

9. Is the following sentence true or false? The empirical formula of a compound

is always the same as the molecular formula. _____false_____

10. Look at Figure 10.16 and Table 10.3. Name three compounds that have an empirical formula of CH.

ethyne, styrene, benzene

11. Fill in the labels on the diagram below.

SO_3 molecule   composed of   ___1___ S atom   and   3 ___O___ atoms

MICROSCOPIC INTERPRETATION

$$SO_3$$

MACROSCOPIC INTERPRETATION

1 mol SO_3 composed of   ___6.02×10^{23}___ sulfur atoms
and
___3___ × (___6.02___ × 10^{23}) oxygen atoms

▶ **Molecular Formulas** (pages 311–312)

12. The molecular formula of a compound is either the same as its empirical

formula or a _____ whole-number multiple _____ of it.

13. What do you need to know to calculate the molecular formula of a compound?

You need the empirical formula of the compound and its molar mass.

14. If you divide the molar mass of a compound by the empirical formula mass, what is the result?

The result is the number of empirical formula units in a molecule of the compound.

15. What factor would you use to convert the empirical formula of a compound to a molecular formula?

the number of empirical formula units in a molecule of the compound

Reading Skill Practice

By looking carefully at photographs and illustrations in textbooks, you can better understand what you have read. Look carefully at Figure 10.15 on page 309. What important idea does this illustration communicate?

This illustration shows that a chemical formula, in this example CO_2, can be understood on the microscopic level in terms of atoms or on the macroscopic level in terms of moles of atoms.

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GUIDED PRACTICE PROBLEMS

GUIDED PRACTICE PROBLEM 1 (page 289)

1. If 0.20 bushels is 1 dozen apples and a dozen apples has a mass of 2.0 kg, what is the mass of 0.50 bushel of apples?

Analyze

Step 1. List the knowns and the unknown.
Knowns

number of bushels = 0.5 bushel

0.20 bushel = 1 dozen apples

1 dozen apples = 2.0 kg

Unknown

Mass of 0.5 bushel of apples = ? kg

Use dimensional analysis to convert the number of bushels to the mass of apples, by following this sequence of conversions:

Number of bushels → dozens of apples → mass of apples

Calculate

Step 2. Solve for the unknown.
The first conversion factor is: $\dfrac{1\,\text{dozen apples}}{0.20\,\text{bushel}}$

The second conversion factor is: $\dfrac{2.0\,\text{kg apples}}{1\,\text{dozen apples}}$

Multiplying the number of bushels by these two conversion factors gives the answer in kilograms.

mass of apples = $0.50\,\cancel{\text{bushels}} \times \dfrac{1\,\cancel{\text{dozen apples}}}{0.20\,\cancel{\text{bushel}}} \times \dfrac{2.0\,\text{kg apples}}{1\,\cancel{\text{dozen apples}}}$ = 5.0 kg apples

The mass of 0.50 bushel of apples is _____.

Evaluate

Step 3. Does the result make sense?

Because a dozen apples is 2.0 kg and 0.5 bushels is more than two dozen but less

than three dozen, the mass should be more than 4 kg (2 dozen $\times$ 2.0 kg) and less

than 6 kg (3 dozen $\times$ 2.0 kg).

GUIDED PRACTICE PROBLEM 3 (page 291)

3. How many moles is 2.80×10^{24} atoms of silicon?

Step 1. List what you know.	2.80×10^{24} atoms of Si	
	$\boxed{6.02 \times 10^{23}}$ atoms in one mole	

Step 2. Multiply the atoms of silicon by a mol/atoms conversion factor.

$$2.80 \times 10^{24} \text{ atoms Si} \times \frac{1 \text{ mol}}{\boxed{6.02 \times 10^{23}} \text{ atoms Si}}$$

Step 3. Divide.　　　$\boxed{4.65}$ mol

GUIDED PRACTICE PROBLEM 5 (page 292)

5. How many atoms are in 1.14 mol SO_3?

Analyze

Step 1. List the knowns and the unknown.
Knowns

number of moles = 1.14 mol SO_3

$1 \text{ mol SO}_3 = 6.02 \times 10^{23}$ molecules SO_3

$1 \text{ molecule SO}_3 = 4$ atoms (1 S atom and 3 O atoms)

Unknown

$1.14 \text{ mol SO}_3 = ?$ atoms

Calculate

Step 2. Solve for the unknown.
The first conversion factor is $\dfrac{6.02 \times 10^{23} \text{ molecules of water}}{1 \text{ mol water}}$

The second conversion factor is $\dfrac{4 \text{ atoms}}{1 \text{ molecule SO}_3}$.

Multiply moles of SO_3 by these conversion factors:

number of atoms = $1.14 \text{ mol SO}_3 \times \dfrac{6.02 \times 10^{23} \text{ molecules of SO}_3}{1 \text{ mol SO}_3} \times \dfrac{4 \text{ atoms}}{1 \text{ molecule SO}_3}$.

= $\underline{\quad 2.75 \times 10^{24} \text{ atoms} \quad}$

Evaluate

Step 3. Does the result make sense?

Because 4 atoms are in a molecule of SO_3 and there is a little more than one mole of

molecules, the answer should be more than 4 times Avogadro's number of atoms.

CHAPTER 10, Chemical Quantities *(continued)*

EXTRA PRACTICE (similar to Practice Problem 5, page 292)

4. How many molecules is 0.360 mol of water?

Analyze

Step 1. List the knowns and the unknown.
Knowns

Number of mols of water = 0.360 mol water

1 mol water = 6.02×10^{23} molecules water

Unknown

0.360 mol water = ? molecules water

Calculate

Step 2. Solve for the unknown.
The conversion factor is $\dfrac{6.02 \times 10^{23} \text{ molecules of water}}{1 \text{ mol water}}$

Multiplying mols of water by this conversion factor will give the answer

molecules of water = 0.360 ~~mol water~~ $\times \dfrac{6.02 \times 10^{23} \text{ molecules of water}}{1 \text{ ~~mol water~~}}$

= $\underline{\quad 2.17 \times 10^{23} \text{ molecules of water} \quad}$

Evaluate

Step 3. Does the result make sense?

Since the given number of mols of water is about one-third mol, the result should

be about one-third of Avogadro's number of molecules ($1/3 \times 6 = 2$).

GUIDED PRACTICE PROBLEM 7 (page 296)

7. Find the molar mass of PCl_3.

Analyze

Step 1. List the knowns and the unknown.
Knowns

Molecular formula = PCl_3

1 molar mass P = 31.0 g P

1 molar mass Cl = 35.5 g Cl

Unknown

molar mass PCl_3 = ? g

Calculate

Step 2. Solve for the unknown.
Convert moles of phosphorus and chlorine to grams of phosphorus and chlorine. Then add to get the results.

$$1 \text{ mol P} \times \frac{31.0 \text{ g P}}{1 \text{ mol P}} = 31.0 \text{ g P}$$

$$3 \text{ mol Cl} \times \frac{35.5 \text{ g Cl}}{1 \text{ mol Cl}} = 106.5 \text{ g Cl}$$

molar mass of PCl_3 = _____137.5 g_____

Evaluate

Step 3. Does the result make sense?

The answer is the sum of the molar mass of phosphorus and three times the molar

mass of chlorine, expressed to the tenths decimal place.

EXTRA PRACTICE (similar to Practice Problem 5, page 292)

5. How many atoms are there in 2.00 moles of SO_3?

$$2.00 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules } SO_3}{\text{mol}} \times \frac{4 \text{ atoms}}{\text{molecule } SO_3} = 4.82 \times 10^{24} \text{ atoms}$$

EXTRA PRACTICE (similar to Practice Problem 7, page 296)

7. Find the molar mass of table salt (sodium chloride).

The formula for sodium chloride is NaCl.

molar mass NaCl = 1 mol Na + 1 mol Cl

$$= (1 \text{ mol Na} \times \frac{23.0 \text{ g}}{1 \text{ mol Na}}) + (1 \text{ mol Cl} \times \frac{35.5 \text{ g}}{1 \text{ mol Cl}})$$

$$= 23.0 \text{ g} + 35.5 \text{ g} = 58.5 \text{ g}$$

EXTRA PRACTICE (similar to Practice Problem 8, page 296)

8. What is the mass of 1 mole of ozone (O_3)?

$$1 \text{ mol } O_3 \times \frac{48.0 \text{ g } O_3}{1 \text{ mol } O_3} = 48.0 \text{ g } O_3$$

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GUIDED PRACTICE PROBLEM 16 (page 298)

16. Find the mass, in grams, of 4.52×10^{-3} mol $C_{20}H_{42}$.

Analyze

Step 1. List the known and the unknown.

Known

number of moles $= 4.52 \times 10^{-3}$ mol $C_{20}H_{42}$

Unknown

mass $= ?$ g $C_{20}H_{42}$

Calculate

Step 2. Solve for the unknown.
Determine the molar mass of $C_{20}H_{42}$:

1 mol $C_{20}H_{42} = 20 \times 12.0$ g $+ 42 \times 1.0$ g $= 282$ g

Multiply the given number of moles by the conversion factor:

mass $= 4.52 \times 10^{-3}$ ~~mol $C_{20}H_{42}$~~ $\times \dfrac{282 \text{ g } C_{20}H_{42}}{1 \text{ mol } C_{20}H_{42b}} = $ _____1.27 g $C_{20}H_{42}$_____

Evaluate

Step 3. Does the result make sense?

The amount of substance is a little more than four-one thousandths of a mole, so

the mass should be only a small fraction of the molar mass.

EXTRA PRACTICE (similar to Practice Problem 17, page 298)

17. Calculate the mass, in grams, of 10 mol of sodium sulfate (Na_2SO_4).

$10 \times (2$ mol Na $+ 1$ mol S $+ 4$ mol O$)$

$= 10 \times (2$ ~~mol Na~~ $\times \dfrac{23.0 \text{ g}}{\text{mol Na}} + 1$ ~~mol S~~ $\times \dfrac{32.1 \text{ g}}{\text{mol S}} + 4$ ~~mol O~~ $\times \dfrac{16.0 \text{ g}}{\text{mol O}})$

$= 10 \times (46.0$ g $+ 32.1$ g $+ 64.0$ g$)$

$= 10 \times (142.1$ g$) = 1421$ g

Calculate the mass, in grams, of 10 mol of iron(II) hydroxide ($Fe(OH)_2$).

$10 \times (1$ mol Fe $+ 2$ mol O $+ 2$ mol H$)$

$= 10 \times (1$ ~~mol Fe~~ $\times \dfrac{55.8 \text{ g}}{\text{mol Fe}} + 2$ ~~mol O~~ $\times \dfrac{16.0 \text{ g}}{\text{mol O}} + 2$ ~~mol H~~ $\times \dfrac{1.0 \text{ g}}{\text{mol H}})$

$= 10 \times (55.8$ g $+ 32.0$ g $+ 2.0$ g$)$

$= 10 \times 89.8$ g $= 898$ g

GUIDED PRACTICE PROBLEM 18 (page 299)

18. Find the number of moles in 3.70×10^{-1} g of boron.

Analyze

Step 1. List the known and the unknown.
Known

mass = 3.70×10^{-1} g boron

Unknown

number of moles = ? mol boron

The unknown number of moles is calculated by converting the known mass to the number of moles using a conversion factor of mass → moles.

Calculate

Step 2. Solve for the unknown.
Determine the molar mass of boron: 1 mol B = 10.8 g B

Multiply the given mass by the conversion factor relating mass of boron to moles of boron:

$$\text{mass} = 3.70 \times 10^{-1} \; \cancel{\text{g B}} \times \frac{1 \text{ mol B}}{10.8 \; \cancel{\text{g B}}}$$

$$= \underline{\quad 3.43 \times 10^{-2} \text{ mol B} \quad}$$

Evaluate

Step 3. Does the result make sense?

Because the value of the conversion factor is about one-tenth, the numerical result

should be about one-tenth of the given number of grams.

GUIDED PRACTICE PROBLEM 20 (page 301)

20. What is the volume of these gases at STP?
a. 3.20×10^{-3} mol CO_2
b. 3.70 mol N_2

a. 3.20×10^{-3} mol CO_2
Analyze

Step 1. List the knowns and the unknown.
Knowns

number of moles = 3.20×10^{-3} mol CO_2

1 mol CO_2 = 22.4 L CO_2

CHAPTER 10, Chemical Quantities *(continued)*

Unknown

volume = ? L CO_2

To convert moles to liters, use the relationship 1 mol CO_2 = 22.4 L CO_2 (at STP).

Calculate

Step 2. Solve for the unknown.
Multiply the given number of moles of CO_2 by the conversion factor to give the result.

volume = 3.20×10^{-3} ~~mol CO_2~~ $\times$ $\dfrac{22.4 \text{ L } CO_2}{1 \text{ } \cancel{\text{mol }CO_2}}$

= $\underline{\qquad 7.17 \times 10^{-2} \text{ L } CO_2 \qquad}$

Evaluate

Step 3. Does the result make sense?

Because a mole of gas occupies a volume of a little more than 20 liters, the result

should be a little larger than twenty times the given number of moles.

b. 3.70 mol N_2

Analyze

Step 1. List the knowns and the unknown.
Knowns

number of moles = 3.70 mol N_2

1 mol N_2 = 22.4 L N_2

Unknown

volume = ? L N_2

Use the relationship 1 mol N_2 = 22.4 L N_2 (at STP) to convert moles to liters.

Calculate

Step 2. Solve for the unknown.
Multiply the given number of moles of N_2 by the conversion factor to give the result.

volume = 3.70 ~~mol N_2~~ $\times$ $\dfrac{22.4 \text{ L } N_2}{1 \text{ } \cancel{\text{mol }N_2}}$

= $\underline{\qquad 82.9 \text{ L } N_2 \qquad}$

Evaluate

Step 3. Does the result make sense?

Because the number of moles is slightly less than four, the result should be close

to, but less than 88 L.

GUIDED PRACTICE PROBLEM 22 (page 302)

22. A gaseous compound composed of sulfur and oxygen, which is linked to the formation of acid rain, has a density of 3.58 g/L at STP. What is the molar mass of this gas?

Analyze

Step 1. List the knowns and the unknown.
Knowns

density = 3.58 g/L

1 mol (gas at STP) = 22.4 L

Unknown

molar mass = ? g

To convert density (g/L) to molar mass (g/mol), a conversion factor of L/mol is needed.

Calculate

Step 2. Solve for the unknown.
Multiply the density by the conversion factor relating liters and moles.

$$\text{molar mass} = \frac{3.58 \text{ g}}{1 \text{ L}} \times \frac{22.4 \text{ L}}{1 \text{ mol}}$$

$$= \underline{\qquad 80.2 \text{ g/mol} \qquad}$$

Evaluate

Step 3. Does the result make sense?

Multiplying approximately 4 grams/liter by approximately 20 liters/mole yields

about 80 grams/mole.

GUIDED PRACTICE PROBLEM 32 (page 306)

32. A compound is formed when 9.03 g Mg combines with 3.48 g N. What is the percent composition of this compound?

Analyze

Step 1. List the known and the unknowns.

Knowns

mass of Mg = 9.03 g Mg

mass of N = 3.48 g N

mass of compound = 9.03 g Mg + 3.48 g N = 12.51 g

Unknowns

percent Mg = ? %

percent N = ? %

The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To be expressed as a percentage, the ratio must be multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent Mg} = \frac{9.03 \text{ g Mg}}{12.51 \text{ g compound}} \times 100\% = \underline{\quad 72.2\% \text{ Mg} \quad}$$

$$\text{percent N} = \frac{3.48 \text{ g N}}{12.51 \text{ g compound}} \times 100\% = \underline{\quad 27.8\% \text{ N} \quad}$$

Evaluate

Step 3. Does the result make sense?

The percents of the elements of the compound add up to 100%.

72.2% + 27.8% = 100%

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GUIDED PRACTICE PROBLEM 34 (page 307)

34. Calculate the percent composition of these compounds.
  a. ethane (C_2H_6)
  b. sodium bisulfate ($NaHSO_4$)

a. ethane (C_2H_6)

Analyze

Step 1. List the knowns and the unknowns.

Knowns

mass of C in one mole ethane = 2×12.0 g = 24.0 g

mass of H in one mole ethane = 6×1.0 g = 6.0g

molar mass of C_2H_6 = 24.0 g + 6.0 g = 30.0 g

Unknowns

percent C = ? %

percent H = ? %

Because no masses are given, the percent composition can be determined based on the molar mass of the substance. The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To express the ratio as a percent, the ratio is multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent C} = \frac{24.0 \text{ g C}}{30.0 \text{ g compound}} \times 100\% = \underline{\quad 80.0\% \text{ C} \quad}$$

$$\text{percent H} = \frac{6.0 \text{ g H}}{30.0 \text{ g compound}} \times 100\% = \underline{\quad 20.0\% \text{ H} \quad}$$

Evaluate

Step 3. Does the result make sense?

The percents of the elements of the compound add up to 100%.

80.0% + 20.0% = 100%

CHAPTER 10, Chemical Quantities *(continued)*

b. sodium bisulfate ($NaHSO_4$)
Analyze

Step 1. List the knowns and the unknowns.
Knowns

mass of Na in one mole sodium bisulfate = $1 \times 23.0g = 23.0$ g

mass of H in one mole sodium bisulfate = 1×1.0 g = 1.0 g

mass of S in one mole sodium bisulfate = 1×32.1 g = 32.1 g

mass of O in one mole sodium bisulfate = 4×16.0 g = 64.0 g

molar mass of $NaHSO_4$ = 23.0 g + 1.0 g + 32.1 g + 64.0 g = 120.1 g

Unknowns

percent Na = ? %

percent H = ? %

percent S = ? %

percent O = ? %

Because no masses are given, the percent composition can be determined based on the molar mass of the substance. The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To express the ratio as a percent, the ratio is multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent Na} = \frac{23.0 \text{ g Na}}{120.1 \text{ g compound}} \times 100\% = \underline{\quad 19.2\% \text{ Na} \quad}$$

$$\text{percent H} = \frac{1.0 \text{ g H}}{120.1 \text{ g compound}} \times 100\% = \underline{\quad 0.83\% \text{ H} \quad}$$

$$\text{percent S} = \frac{32.1 \text{ g S}}{120.1 \text{ g compound}} \times 100\% = \underline{\quad 26.7\% \text{ C} \quad}$$

$$\text{percent O} = \frac{64.0 \text{ g O}}{120.1 \text{ g compound}} \times 100\% = \underline{\quad 53.3\% \text{ O} \quad}$$

Evaluate

Step 3. Does the result make sense?

The percents of the elements of the compound add up to 100%.

19.2% + 0.83% + 26.7% + 53.3% = 100%

GUIDED PRACTICE PROBLEM 36 (page 310)

36. Calculate the empirical formula of each compound.
    a. 94.1% O, 5.9% H
    b. 67.6% Hg, 10.8% S, 21.6% O

a. 94.1% O, 5.9% H

Analyze

Step 1. List the knowns and the unknown.
Knowns

percent composition: 94.1% O, 5.9% H

molar mass O = 16.0 g/mol O

molar mass H = 1.0 g/mol H

Unknown

empirical formula = $H_?O_?$

Use the percent composition to convert to mass, recalling that percent means parts per hundred. Then use the molar mass to convert to number of moles. Finally, determine whole number ratios based on the number of moles of each element per 100 grams of compound.

Calculate

Step 2. Solve for the unknown.
One hundred grams of compound has 5.9 g H and 94.1 g O.
Multiply by conversion factors relating moles of the elements to grams.

$$5.9 \; \cancel{g \, H} \times \frac{1 \; mol \; H}{1.0 \; \cancel{g \, H}} = 5.9 \; mol \; H$$

$$94.1 \; \cancel{g \, O} \times \frac{1 \; mol \; O}{16.0 \; \cancel{g \, O}} = 5.88 \; mol \; O$$

So the mole ratio for 100 g of the compound is $H_{5.9}O_{5.9}$. But formulas must have whole number subscripts. Divide each molar quantity by the smaller number of moles. This will give 1 mol for the element with the smaller number of moles. In this case, the ratio is one-to-one and so the empirical formula is simply H_1O_1. However, a subscript of one is never written, so the answer is __HO__.

Evaluate

Step 3. Does the result make sense?

The subscripts are whole numbers and the percent composition of this empirical

formula equals the percent compostion given in the original problem.

CHAPTER 10, Chemical Quantities *(continued)*

b. 67.6% Hg, 10.8% S, 21.6% O

Analyze

Step 1. List the knowns and the unknown.

Knowns

percent composition: 67.6% Hg, 10.8% S, 21.6% O

molar mass Hg = 200.6 g/mol Hg

molar mass S = 32.1 g/mol S

molar mass O = 16.0 g/mol O

Unknown

empirical formula = $Hg_?S_?O_?$

Use the percent composition to convert to mass. Then use molar mass to convert to number of moles. Finally, determine whole number ratios based on the number of moles of each element per 100 grams of compound.

Calculate

Step 2. Solve for the unknown.
One hundred grams of compound has 67.6 g Hg, 10.8 g S, and 21.6 g O.
Multiply by a conversion factor relating moles to grams.

$$67.6 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} = 0.337 \text{ mol Hg}$$

$$10.8 \text{ g S} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} = 0.336 \text{ mol S}$$

$$21.6 \text{ g O} \times \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 1.35 \text{ mol O}$$

So the mole ratio for 100 g of the compound is $Hg_{0.34}S_{0.34}O_{1.35}$.
Divide each molar quantity by the smaller number of moles.

$$\frac{0.34 \text{ mol Hg}}{0.34} = 1 \text{ mol Hg}$$

$$\frac{0.34 \text{ mol S}}{0.34} = 1 \text{ mol S}$$

$$\frac{1.35 \text{ mol O}}{1.35} = 4 \text{ mol O}$$

The empirical formula is _____ $HgSO_4$ _____ .

Evaluate

Step 3. Does the result make sense?

The subscripts are whole numbers and the percent composition of this empirical

formula equals the percent composition given in the original problem.

GUIDED PRACTICE PROBLEM 38 (page 312)

38. Find the molecular formula of ethylene glycol, which is used as antifreeze. The molar mass is 62 g/mol and the empirical formula is CH_3O.

Analyze

Step 1. List the knowns and the unknown.
Knowns

molar mass = 62 g/mol

empirical formula = CH_3O

Unknown

molecular formula = $C_?H_?O_?$

Calculate

Step 2. Solve for the unknown.
First, calculate the empirical formula mass (efm):

$$1 \; \cancel{mol \, C} \times \frac{12 \text{ g C}}{1 \; \cancel{mol \, C}} = 12 \text{ g C}$$

$$3 \; \cancel{mol \, H} \times \frac{1.0 \text{ g H}}{1 \; \cancel{mol \, H}} = 3 \text{ g H}$$

$$1 \; \cancel{mol \, O} \times \frac{16 \text{ g O}}{1 \; \cancel{mol \, O}} = 16 \text{ g O}$$

So efm = 12g + 3 g + 16 g = 31 g.

Divide the molar mass by the empirical formula mass:
Molar mass/efm = 62 g/31 g = 2

Multiply subscripts in the empirical formula by this value.

The molecular formula is _____ $C_2H_6O_2$ _____ .

Evaluate

Step 3. Does the result make sense?

The molecular formula has the given molar mass, and can be reduced to the

empirical formula.