SECTION 4.2 QUESTIONS

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Understanding Concepts

1. 9×12 u = 108 u (assume *both* numbers are exact), which is very close to the average value for silver, Ag, which is actually 107.87 u.
2. The atomic mass unit is defined as 1/12 the mass of a C-12 atom. Atomic mass is the actual average mass of the atoms of a given element.
3. Because the elements are compared to each other based on the way they combine when they react, the reaction proportions must be known.
4. Chlorine atoms have an average atomic mass of 35.45 u, because chlorine is made up of about 75% Cl-35 atoms, and about 25% Cl-37 atoms.
5. Assume 1000 B atoms, for convenience
   19.8% or 198 atoms are B-10
   80.2% or 802 atoms are B-11

$$m_{tot} = (198 \times 10\ \text{u}) + (802 \times 11\ \text{u})$$

$$m_{tot} = 10\ 802\ \text{u}$$

$$m_{av} = \frac{10\ 802\ \text{u}}{1000}$$

$$m_{av} = 10.8\ \text{u}$$

   The average atomic mass for boron is 10.8 u, calculated by this method.

4.3 THE MOLE AND MOLAR MASS

PRACTICE

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Understanding Concepts

1. A mole of anything is the same number of things as the number of C atoms there would be in exactly 12 g of the isotope C-12.
2. Avogadro's constant is 6.02×10^{23} (rounded to three digits ...).
   Note: Since the concepts of "pure" C-12 and "exactly" 12 g are imaginary, there is no pretense in the scientific community that we will ever know the "exact" value for Avogadro's constant. The mole is a purely theoretical definition. As technology improves, we are, of course, able to determine the value to greater precision. Rounded to six digits, the precision routinely stated in postsecondary level work, the currently accepted value is $6.022\ 14 \times 10^{23}$. The Canadian Metric Practice Guide lists 8 digits — $6.022\ 136\ 7 \times 10^{23}$. The most precise recent reported value, obtained from ion X-ray diffraction evidence, is $6.022\ 141\ 99 \times 10^{23}$. This constant, like many others that are frequently used, is usually rounded to three digits for high-school calculations.
3. 602 000 000 000 000 000 000 000 (only the first 3 digits are significant)
   Note: It is useful for students to extend this number to see how large it is, being mindful that all those trailing zeros represent numbers they don't know. It is also thought-provoking for students to consider that when written in scientific notation to the usual high-school text precision, the error of the number is 2×10^{20}, meaning we blithely ignore some two hundred millions of millions of millions, as being too small to make a noticeable difference.
4. $$N_{CO_2} = 3.00\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}$$

$$N_{CO_2} = 1.81 \times 10^{24}\ \text{molecules}$$

   There are 1.81×10^{24} molecules of carbon dioxide in the sample.

5. $N_{Ar} = 0.500 \cancel{\text{mol}} \times \dfrac{6.02 \times 10^{23} \text{ atoms}}{1 \cancel{\text{mol}}}$

$N_{Ar} = 3.01 \times 10^{23}$ atoms

There are 3.01×10^{23} atoms of argon in the sample.

6. (a) $m = \dfrac{1.43 \text{ kg}}{12 \text{ oranges}}$ (12 is an exact value — counted)

$m = 0.119$ kg/orange = 119 g/orange

The average mass of one of these oranges is 119 g.

(b) $m = \dfrac{1.01 \text{ g}}{1 \cancel{\text{mol}}} \times \dfrac{1 \cancel{\text{mol}}}{6.02 \times 10^{23} \text{ H atoms}}$

$m = 1.68 \times 10^{-24}$ g/H atom (average value)

The average mass of an atom in a hydrogen sample is 1.68×10^{-24} g.

Note: No actual hydrogen atom, of course, really has this mass value, since the relative atomic molar mass used to derive it is an *average* value, itself derived from the natural isotopic composition of the element.

PRACTICE

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Understanding Concepts

7. Molar mass is the mass of one mole (the Avogadro number) of any entity, usually atoms, molecules, ions, or ionic compound formula units. The SI unit is technically kg/mol, but is commonly used and stated as g/mol.

8. 1 mol $Ca(OH)_2$ = 1 mol Ca + 2 mol O + 2 mol H

$M_{Ca(OH)_2} = [(40.08 \times 1) + (16.00 \times 2) + (1.01 \times 2)]$ g/mol

$M_{Ca(OH)_2} = 74.10$ g/mol

The molar mass of calcium hydroxide is 74.10 g/mol.

9. 1 mol Cl_2 = 2 mol Cl

$M_{Cl_2} = (35.45 \times 2)$ g/mol

$M_{Cl_2} = 70.90$ g/mol

The molar mass of chlorine is 70.90 g/mol

10. 1 mol OH^- = 1 mol O + 1 mol H + 1 mol e^-

$M_{OH^-} = [(16.00 \times 1) + (1.01 \times 1) + (\text{negligible mass})]$ g/mol

$M_{OH^-} = 17.01$ g/mol

The molar mass of hydroxide ions is 17.01 g/mol.

Note: Students are taught routinely to simply ignore the electrical charges of ions when calculating molar masses, because the tiny variation in mass due to electron increase/decrease is negligible, never enough to affect the value.

11. The element is most likely gold, $Au_{(s)}$, which has a molar mass listed on the periodic table of elements of 196.97 g/mol.

12. $m = 8.0 \cancel{\text{mol}} \times \dfrac{67.2 \text{ g}}{\cancel{\text{mol}}}$

$m = 5.4 \times 10^2$ g = 0.54 kg

The mass of the substance is 0.54 kg.

13. (a) Diatomic means composed of two-atom molecules.
(b) Hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine form diatomic molecules.
(c) Avogadro's number of molecules (6.02×10^{23} molecules) are present.

SECTION 4.3 QUESTIONS

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Understanding Concepts

1. Avogadro's constant is the number of atoms of carbon in exactly 12 g of the isotope C-12, also called one mole. It is useful in chemistry because a mole is defined so that the mass of this amount in grams will equal numerically the relative atomic mass of any atom.
2. Atomic mass is the average value of the mass of atoms of a particular element. Molar mass is the mass of one mole of any entity.
3. 1 mol $C_{12}H_{22}O_{11}$ = 12 mol C + 22 mol H + 11 mol O

$M_{C_{12}H_{22}O_{11}} = [(12.01 \times 12) + (1.01 \times 22) + (16.00 \times 11)]$ g/mol

$M_{C_{12}H_{22}O_{11}} = 342.34$ g/mol

$N_A = 6.02 \times 10^{23}$ entities/mol

$m_{C_{12}H_{22}O_{11}} = ?$

$$m = \frac{342.34 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{6.02 \times 10^{23} \text{ molecules}}$$

$$m = 5.69 \times 10^{-22} \text{ g/molecule (on average)}$$

The average mass of a molecule in a sucrose sample is 5.69×10^{-22} g.

Note: No actual single sucrose molecule, of course, has this mass value. It is an average value derived from the natural isotope mixtures of elements in this compound.

4. 1 mol C_8H_{18} = 8 mol C + 18 mol H

$M_{C_8H_{18}} = [(12.01 \times 8) + (1.01 \times 18)]$ g/mol

$M_{C_8H_{18}} = 114.26$ g/mol

The molar mass of octane is 114.26 g/mol.

Note: You may find it convenient, from this point on in the course, to assume that students will calculate molar masses accurately, i.e., not requiring them to show the addition of periodic table values. Since all of the periodic table (average atomic molar mass) values are given to two decimal places, and the precision rule for addition always applies, any molar mass calculation will automatically be precise to hundredths of a gram per mole.

Making Connections

5. Molar masses:

	$H_{2(g)}$	2.02 g/mol
	$He_{(g)}$	4.00 g/mol
	$N_{2(g)}$	28.02 g/mol
	$O_{2(g)}$	16.00 g/mol
	$CO_{2(g)}$	44.01 g/mol

The density of each gas is found as follows; assume the volume of a mole of each is approximately 22.4 L.

$$\text{density}_{H_2} = \frac{2.02 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$$\text{density}_{H_2} = 0.0902 \text{ g/L}$$

The density of hydrogen gas at STP is 0.0902 g/L.

$$\text{density}_{He} = \frac{4.00 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{He} = 0.179 \text{ g/L}$

The density of helium gas at STP is 0.179 g/L.

$$\text{density}_{N_2} = \frac{28.02 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{N_2} = 1.25 \text{ g/L}$

The density of nitrogen gas at STP is 1.25 g/L.

$$\text{density}_{O_2} = \frac{32.00 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{O_2} = 1.43 \text{ g/L}$

The density of oxygen gas at STP is 1.43 g/L.

$$\text{density}_{CO_2} = \frac{44.01 \text{ g}}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{22.4 \text{ L}}$$

$\text{density}_{CO_2} = 1.96 \text{ g/L}$

The density of carbon dioxide gas at STP is 1.96 g/L.

Note : The molar volume of 22.4 L/mol at STP for gases is an approximation derived from a theoretical "ideal" gas system, rounded to three significant digits. See page 469 of the text (Did You Know?) for more precise values for some real gases.

Reflecting

6. Molar mass is a conversion factor that can be used to convert measured masses of substances into numerical amounts, in moles. This will be useful because the numbers in chemical reaction equations represent numerical values.

4.4 CALCULATIONS INVOLVING THE MOLE CONCEPT

PRACTICE

(Page 172)

Understanding Concepts

1. $m_{NaCl} = 2.5 \text{ g}$

$M_{NaCl} = 58.44 \text{ g/mol}$

$n_{NaCl} = ?$

$$n_{NaCl} = 2.5 \cancel{\text{g}} \times \frac{1 \text{ mol}}{58.44 \cancel{\text{g}}}$$

$$n_{NaCl} = 0.043 \text{ mol} = 43 \text{ mmol}$$

The amount of sodium chloride is 43 mmol.

2. $m_{C_6H_{12}O_6} = 1.0 \text{ kg}$

$M_{C_6H_{12}O_6} = 180.18 \text{ g/mol}$

$n_{C_6H_{12}O_6} = ?$

$$n_{C_6H_{12}O_6} = 1.0 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{180.18 \cancel{\text{g}}}$$

$$n_{C_6H_{12}O_6} = 0.056 \text{ kmol} = 5.6 \text{ mol}$$

The amount of glucose is 5.6 mol.