# Molar Mass



Hydrogen is an important industrial chemical. The most common production method is by the reaction of a hydrocarbon with steam (**Figure 1**). The chemical equation for this reaction, using methane, is

$$CH_4(g) + H_2O(g) \rightarrow 3\ H_2(g) + CO(g)$$

This chemical equation tells you that 1 molecule of water reacts with every molecule of methane. Hydrogen manufacturers must have a convenient method of measuring exactly the required quantities of both reactants. Simply combining identical masses of the two substances will not work because the atomic masses of carbon, hydrogen, and oxygen are different. Therefore, 1 kg of methane contains a different number of molecules than 1 kg of water. To get around this problem, chemical manufacturers calculate the amount of each substance required (in moles). To do this, they must first determine the molar mass of each substance. The **molar mass** is the mass in grams of 1 mole or $6.02 \times 10^{23}$ entities of that substance. The units of molar mass are g/mol. Chemists often use the symbol $M$ to represent molar mass, with the chemical symbol or formula as a subscript. The symbol for the molar mass of water is therefore $M_{H_2O}$.

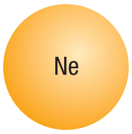
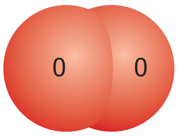
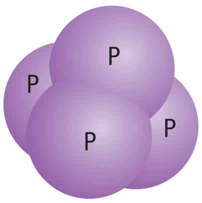
**Figure 1** A researcher studies how to use steam to produce hydrogen gas from hydrocarbons. This process is called steam-reforming.

**molar mass** the mass of 1 mol of a substance; unit symbol g/mol

## Molar Mass of Elements

The molar mass of a monatomic element such as neon is equal to the mass given on the periodic table. The molar mass of neon is therefore 20.18 g/mol. The molar mass of a molecular element such as oxygen is determined by multiplying the molar mass of the element by the number of atoms per molecule (**Table 1**).

**Table 1** Calculating Molar Masses of Elements

| Name and formula | Model | Molar mass calculation |
|---|---|---|
| neon, Ne |   | $M_{Ne} = 1 \times 20.18\ \frac{g}{mol} = 20.18\ \frac{g}{mol}$ |
| oxygen, $O_2$ |   | $M_{O_2} = 2 \times 16.00\ \frac{g}{mol} = 32.00\ \frac{g}{mol}$ |
| phosphorus, $P_4$ |   | $M_{P_4} = 4 \times 30.97\ \frac{g}{mol} = 123.9\ \frac{g}{mol}$ |

**LEARNING TIP**

**Significant Digits and Counting Numbers**

Some values are exact. For example, there are exactly 2 atoms of oxygen in a molecule of $O_2$. In calculations such as

$$2 \times 16.00\ \frac{g}{mol} = 32.00\ \frac{g}{mol}$$

you should always treat a "counting number" as if it has an infinite number of significant digits. The number of significant digits in the answer is determined by the measured quantity in the calculation. In this case, 16.00 g/mol has 4 significant digits so the answer should have 4 significant digits.

### Numerical Equivalency of Atomic and Molar Masses

The mass of 1 mole of a monatomic element, expressed in g/mol, has the same numerical value as the average atomic mass of the element expressed in atomic mass units (u). How convenient! This means that, on the periodic table, only one mass value need be given for each element (**Figure 2**). This is true because of the value of Avogadro's constant. We can use carbon to illustrate this. According to the periodic

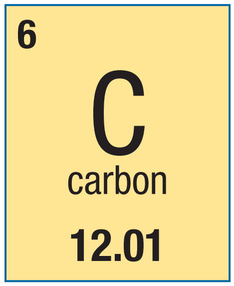



**Figure 2** The atomic mass of a carbon atom, 12.01 u, has the same numerical value as the mass of 1 mole of carbon, 12.01 g.

**LEARNING TIP**

**Significant Digits**

If we want a fairly precise answer—say to 4 significant digits—we need to use values in the calculation with at least this many digits. Therefore, in this calculation we are using Avogadro's constant with 3 decimal places ($6.022 \times 10^{23}$), instead of the more usual 2 decimal places.

table, 1 carbon atom has an average atomic mass of 12.01 u. Chemists define one atomic mass unit (1 u) as follows:

$$1\ \text{u} = 1.661 \times 10^{-24}\ \text{g}$$

The mass of one carbon atom in grams is therefore

$$(12.01\ \cancel{\text{u}})\left(\frac{1.661 \times 10^{-24}\ \text{g}}{1\ \cancel{\text{u}}}\right) = 1.99486 \times 10^{-23}\ \text{g}\ [\text{extra digits carried}]$$

Recall that Avogadro's constant is the number of entities in a mole of a substance. We can multiply the mass of one carbon atom by Avogadro's constant (expressed to 4 significant digits) to find the mass of 1 mol of carbon atoms:

$$\left(\frac{1.99486 \times 10^{-23}\ \text{g}}{1\ \cancel{\text{atom}_\text{c}}}\right)(6.022 \times 10^{23}\cancel{\text{atom}_\text{c}}) = 12.01\ \text{g}$$

The mass of 1 mol of carbon atoms is 12.01 g. The molar mass of carbon is therefore 12.01 g/mol.

As you can see, the molar mass of carbon and the mass of 1 atom of carbon have the same numerical value. Remember, however, that 12.01 g/mol and 12.01 u represent very different quantities of carbon. One mole or 12.01 g of carbon contains $6.022 \times 10^{23}$ atoms—enough to fill one-quarter of a small test tube. However, 12.01 u is the mass of one carbon atom, which is too small to be visible even with the most powerful imaging technology.

## Molar Masses of Compounds

We can also describe quantities of molecular and ionic compounds using moles. The molar mass of a compound is the sum of the molar mass of each entity in the compound. For example, carbon dioxide is a common molecular compound (**Figure 3(a)**). Solid carbon dioxide (dry ice) consists of vast numbers of $CO_2$ molecules. Each molecule contains 1 carbon atom covalently bonded to 2 oxygen atoms. Sodium chloride is a common ionic compound (**Figure 3(b)**). A crystal of sodium chloride does not contain molecules. Instead, it consists of alternating positive sodium ions and negative chloride ions. The chemical formula, NaCl, indicates that there is 1 sodium ion for every chloride ion. In Unit 1 you learned that the simplest ratio of ions in an ionic compound is called a formula unit. NaCl is the formula unit for sodium chloride.

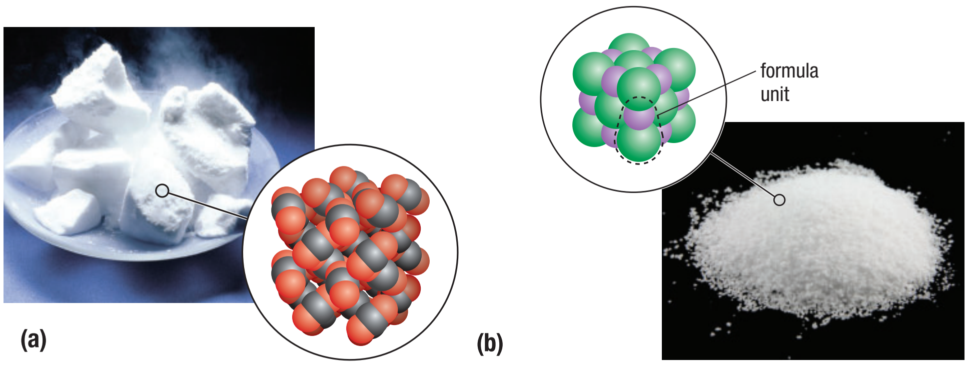



**Figure 3** (a) Molecular solids are made up of molecules, while (b) ionic solids consist of a continuous pattern of alternating positive and negative ions. A formula unit for an ionic compound is the simplest ratio of positive and negative ions in the compound.

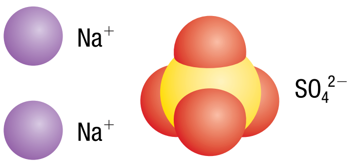
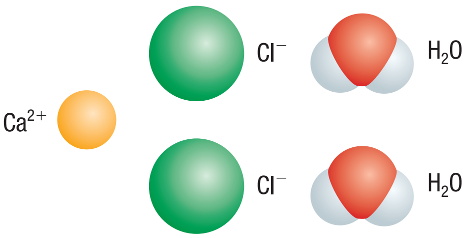
**LEARNING TIP**

**Formula Unit and Molecular Formula**

The formulas for HCl and NaCl look similar, but the entities they represent are quite different. Hydrogen chloride consists of molecules. The atoms within each molecule are joined by covalent bonds. However, there are no separate units of NaCl in a salt crystal. Rather, an ionic compound is a continuous network of positive and negative ions. The formula unit, NaCl, indicates the simplest ratio of these ions in the compound.

**Table 2** shows molar mass calculations for common molecular and ionic compounds. Remember that you have to consider every element in the molecular formula, or formula unit, as you calculate the molar mass.

**Table 2** Common Molecular and Ionic Compounds

| Name and formula | Model | Molar mass calculation |
|---|---|---|
| hydrogen phosphate, $H_3PO_4$ |  | $M_{H_3PO_4} = 3M_H + M_P + 4M_O$<br>$= \left(3 \times 1.01 \frac{g}{mol}\right) + \left(30.97 \frac{g}{mol}\right) + \left(4 \times 16.00 \frac{g}{mol}\right)$<br>$= 3.03 \frac{g}{mol} + 30.97 \frac{g}{mol} + 64.00 \frac{g}{mol}$<br>$M_{H_3PO_4} = 98.00 \frac{g}{mol}$ |
| sodium chloride, NaCl |  | $M_{NaCl} = M_{Na^+} + M_{Cl^-}$<br>$= 22.99 \frac{g}{mol} + 35.45 \frac{g}{mol}$<br>$M_{NaCl} = 58.44 \frac{g}{mol}$ |
| sodium sulfate, $Na_2SO_4$ |  | $M_{Na_2SO_4} = 2M_{Na^+} + M_{SO_4^{2-}}$<br>$= 2M_{Na^+} + M_S + 4M_O$<br>$= \left(2 \times 22.99 \frac{g}{mol}\right) + \left(32.07 \frac{g}{mol}\right) + \left(4 \times 16.00 \frac{g}{mol}\right)$<br>$= 45.98 \frac{g}{mol} + 32.07 \frac{g}{mol} + 64.00 \frac{g}{mol}$<br>$M_{Na_2SO_4} = 142.05 \frac{g}{mol}$ |
| calcium chloride dihydrate, $CaCl_2 \cdot 2H_2O$ |  | $M_{CaCl_2 \cdot 2H_2O} = M_{Ca^{2+}} + 2M_{Cl^-} + 2M_{H_2O}$<br>$= M_{Ca^{2+}} + 2M_{Cl^-} + 2(2M_H + M_O)$<br>$= \left(40.08 \frac{g}{mol}\right) + \left(2 \times 35.45 \frac{g}{mol}\right) +$<br>$2\left(2 \times 1.01 \frac{g}{mol} + 16.00 \frac{g}{mol}\right)$<br>$M_{CaCl_2 \cdot 2H_2O} = 147.02 \frac{g}{mol}$ |

## Tutorial 1 Determining Molar Mass

The molar mass of a monatomic element is provided on the periodic table. Otherwise, determine molar mass by summing the masses of each atom or ion in the chemical formula.

### Sample Problem 1: Calculating the Molar Mass of a Molecular Compound

Calculate the molar mass of carbon dioxide.

**Given:** carbon dioxide, $CO_2$

**Required:** molar mass of $CO_2$, $M_{CO_2}$

**Solution:**

**Step 1.** Look up the molar masses of the elements:

$$M_C = 12.01 \frac{g}{mol}; M_O = 16.00 \frac{g}{mol}$$

**Step 2.** Add the molar masses of the elements, multiplying the molar mass of each element by the number of atoms of that element in the compound. In this case, there is 1 atom of C and 2 atoms of O.

**LEARNING TIP**

**Periodic Table**

You will refer to the periodic table frequently in this Chemistry course. The periodic table is provided on the inside back cover of this textbook and in Appendix B1.

$$M_{CO_2} = M_C + 2M_O$$
$$= \left(12.01\ \frac{g}{mol}\right) + \left(2 \times 16.00\ \frac{g}{mol}\right)$$
$$= 12.01\ \frac{g}{mol} + 32.00\ \frac{g}{mol}$$
$$M_{CO_2} = 44.01\ \frac{g}{mol}$$

**Statement:** The molar mass of carbon dioxide is 44.01 g/mol.

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### Sample Problem 2: Calculating the Molar Mass of an Ionic Compound

Calculate the molar mass of iron(III) oxide.

**Given**: iron(III) oxide, $Fe_2O_3$

**Required:** molar mass of $Fe_2O_3$, $M_{Fe_2O_3}$

**Solution:**

**Step 1.** As in Sample Problem 1, look up the molar masses of the elements.

$$M_{Fe} = 55.85\ \frac{g}{mol}; \quad M_O = 16.00\ \frac{g}{mol}$$

**Step 2.** Add the molar masses of the elements, multiplying the molar mass of each element by the number of atoms of that element in the compound. Note that iron(III) oxide is an ionic compound containing two $Fe^{3+}$ ions and three $O^{2-}$ ions per formula unit. Since the mass of an electron is extremely small, we may assume that the masses of $Fe^{3+}$ ions and $O^{2-}$ ions are the same as the masses of Fe atoms and O atoms, respectively.

$$M_{Fe_2O_3} = 2M_{Fe^{3+}} + 3M_{O^{2-}}$$
$$= 2\left(55.85\ \frac{g}{mol}\right) + 3\left(16.00\ \frac{g}{mol}\right)$$
$$= 111.7\ \frac{g}{mol} + 48.00\ \frac{g}{mol}$$
$$M_{Fe_2O_3} = 159.7\ \frac{g}{mol}$$

**Statement:** The molar mass of iron(III) oxide is 159.7 g/mol.

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### Sample Problem 3: Calculating the Molar Mass of a Hydrate

Calculate the molar mass of iron(III) chloride hexahydrate.

**Given:** iron(III) chloride hexahydrate, $FeCl_3{\cdot}6H_2O$

**Required:** molar mass of $FeCl_3{\cdot}6H_2O$, $M_{FeCl_3 \cdot 6H_2O}$

**Solution:**

**Step 1.** Look up the molar masses of the elements in the compound in the periodic table.

$$M_{Fe^{3+}} = 55.85\ \frac{g}{mol}; M_{Cl^-} = 35.45\ \frac{g}{mol}; M_H = 1.01\ \frac{g}{mol}; M_O = 16.00\ \frac{g}{mol}$$

**Step 2.** Add the molar masses of the elements, multiplying the molar mass of each element by the number of atoms of that element in the compound.

$$M_{FeCl_3 \cdot 6H_2O} = M_{Fe^{3+}} + 3M_{Cl^-} + 6(2M_H + M_O)$$
$$= \left(55.85\ \frac{g}{mol}\right) + 3\left(35.45\ \frac{g}{mol}\right) + 6\left(2 \times 1.01\ \frac{g}{mol} + 16.00\ \frac{g}{mol}\right)$$
$$M_{FeCl_3 \cdot 6H_2O} = 270.32\ \frac{g}{mol}$$

**Statement:** The molar mass of iron(III) chloride hexahydrate is 270.32 g/mol.

## Practice

1. Calculate the molar mass of each of the following substances. Express your answer to 2 decimal places. K/U T/I
   (a) $S_8$ (a component of gunpowder) [ans: 256.56 g/mol]
   (b) $H_2S$ (the smell of rotten eggs) [ans: 34.09 g/mol]
   (c) NaOH (used in paper manufacture) [ans: 40.00 g/mol]
   (d) $Fe(OH)_3$ (a pigment in paint and cosmetics) [ans: 106.88 g/mol]
   (e) $(NH_4)_2S$ (used in textile production) [ans: 68.17 g/mol]
   (f) $Ca_3(PO_4)_2$ (used to make fertilizers) [ans: 310.18 g/mol]
   (g) $MgSO_4{\cdot}7H_2O$ (Epsom salts) [ans: 246.52 g/mol]

## Mini Investigation

### The Mole Exhibit

**Skills:** Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1, A3.2

In this activity, your teacher will ask you to measure 1.00 mol of a pure substance and display it in the classroom.

**Equipment and Materials:** chemical safety goggles; lab apron; balance; scoopula; 250 mL beaker; samples of substances assigned by your teacher

1. Calculate the molar mass of your assigned substance.
2. Once your teacher approves your calculations, put on your chemical safety goggles and lab apron.
3. Use the balance to measure out 1.00 mol of your substance into the beaker.

A. What is the molar mass of your substance? T/I

B. How many entities of your substance are in the beaker? K/U

## Calculations Involving Molar Masses

Chemists routinely use mass to measure the quantity of a chemical needed for an investigation. As such, chemists need a convenient way to determine the amount in moles, $n$, in a given mass in grams, $m$, of a sample. As you will see, molar mass, $M$, is the quantity that mathematically connects the amount of a substance to its mass.

**UNIT TASK BOOKMARK**

You will use the concept of molar mass as you perform the calculations necessary for the Unit Task on page 352.

### Converting Mass to Amount

We can use logic to develop a formula that relates amount, mass, and molar mass. Suppose a chemist needs 4.00 g of sodium hydroxide for an investigation. Earlier, you determined that the mass of one mole of sodium hydroxide is 40.00 g/mol. Therefore, 4.00 g of sodium hydroxide represents one tenth or a mole, or 0.100 mol. How did we arrive at this answer? We divided the mass (4.00 g) by the molar mass (40.00 g/mol):

$$n_{\text{NaOH}} = \frac{4.00\ \text{g}}{40.00\ \dfrac{\text{g}}{\text{mol}}}$$

$$= 4.00\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{40.00\ \cancel{\text{g}}}\right)$$

$$n_{\text{NaOH}} = 0.100\ \text{mol}$$

In general, the mathematical relationship between $m$, $n$ and $M$, therefore, is

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

### Converting Amount to Mass

If sodium hydroxide has a molar mass of 40.00 g, we can predict that 1.50 mol of sodium hydroxide has a mass of 60.0 g. We can get this answer by rearranging the equation $n = \dfrac{m}{M}$ into the form $m = nM$.

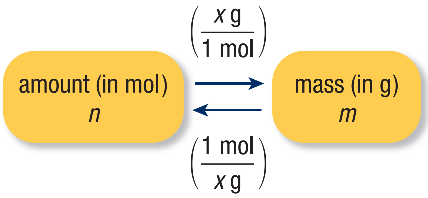


**Figure 4** Molar mass is the mathematical link between the amount in moles and the mass in grams.

**Figure 5** This glass of cola contains 40.0 g of sugar (sucrose).

$$m_{\text{NaOH}} = 1.50\ \cancel{\text{mol}} \times \left(\frac{40.00\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$m_{\text{NaOH}} = 60.0\ \text{g}$$

As you can see from these examples, the mass of a chemical can be converted to an amount, and vice versa, by multiplying by a conversion factor. This factor is either the molar mass (units g/mol) or the reciprocal molar mass (units mol/g) (**Figure 4**).

## Tutorial 2 Converting among Amount, Mass, and Molar Mass

The Sample Problems in this tutorial illustrate how to calculate the mass, *m*, of a sample of a substance given the amount of that substance (in mol), and vice versa.

### **Sample Problem 1:** Calculating Amount from Mass

According to the manufacturer, a typical can of cola contains 40.0 g of sucrose, $C_{12}H_{22}O_{11}$ (**Figure 5**). Calculate the amount of sucrose in 40.0 g.

**Given:** $m_{C_{12}H_{22}O_{11}} = 40.0\ \text{g}$

**Required:** amount of sucrose, $n_{C_{12}H_{22}O_{11}}$

**Solution:**

**Step 1.** Calculate the molar mass of sucrose, $M_{C_{12}H_{22}O_{11}}$.

$$M_{C_{12}H_{22}O_{11}} = 12\left(12.01\ \frac{\text{g}}{\text{mol}}\right) + 22\left(1.01\ \frac{\text{g}}{\text{mol}}\right) + 11\left(16.00\ \frac{\text{g}}{\text{mol}}\right)$$

$$M_{C_{12}H_{22}O_{11}} = 342.34\ \frac{\text{g}}{\text{mol}}$$

**Step 2.** Use the mass of sucrose (given in the question) and the molar mass of sucrose to calculate the amount of sucrose. To do this, you need to develop a conversion factor using the molar mass. The conversion factor is either

$$\frac{1\ \text{mol}}{342.34\ \text{g}} \quad \text{or} \quad \frac{342.34\ \text{g}}{1\ \text{mol}}$$

Since you want to solve for an amount (in mol) of sucrose, multiply by the factor that has mol in the numerator, as follows:

$$n_{C_{12}H_{22}O_{11}} = 40.0\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{342.34\ \cancel{\text{g}}}$$

$$n_{C_{12}H_{22}O_{11}} = 0.117\ \text{mol}$$

**Statement:** There is 0.117 mol of sucrose in a 40.0 sample of sucrose.

### **Sample Problem 2:** Calculating Mass from Amount

A litre of human blood typically contains 4.0 mmol of glucose, $C_6H_{12}O_6$. Calculate the mass of this amount of glucose.

**Given:** $m_{C_6H_{12}O_6} = 40.0\ \text{mmol}$

**Required:** mass of glucose, $M_{C_6H_{12}O_6}$

**Solution:** $4.0\ \text{mmol} = 4.0 \times 10^{-3}\ \text{mol}$

$$M_{C_6H_{12}O_6} = 6\left(12.01\ \frac{\text{g}}{\text{mol}}\right) + 12\left(1.01\ \frac{\text{g}}{\text{mol}}\right) + 6\left(16.00\ \frac{\text{g}}{\text{mol}}\right)$$

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

$$m_{C_6H_{12}O_6} = (4.0 \times 10^{-3}\ \cancel{\text{mol}})\left(\frac{180.18\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$m_{C_6H_{12}O_6} = 0.72\ \text{g}$$

**Statement:** The mass of 4.0 mmol of glucose in a litre of blood is 0.72 g of glucose.

## Practice

SKILLS HANDBOOK A6.3

1. Calculate the amount of pure substance in each of the following samples: T/I
   (a) a 500 g box of table salt (assume pure sodium chloride) [ans: 8.56 mol]
   (b) 14.2 g of aluminum in a typical pop can [ans: 0.526 mol]
   (c) 1.00 kg of calcium oxide, CaO, used to neutralize a soil sample [ans: 17.8 mol]
   (d) 1.75 kg of hydrogen chloride, HCl, in a jug of concentrated hydrochloric acid [ans: 48.0 mol]
   (e) 200.0 mg of ibuprofen, $C_{13}H_{18}O_2$, in a headache medication [ans: $9.694 \times 10^{-4}$ mol]
2. Calculate the mass of the following amounts of a pure substance: T/I
   (a) 0.80 mol of hydrogen peroxide, $H_2O_2$, in a bottle of antiseptic [ans: 27 g]
   (b) 3.25 mol of sodium hydrogen sulfate, $NaHSO_4$, in a container of bathroom cleaner [ans: $3.90 \times 10^2$ g]
   (c) 5.0 mmol of calcium carbonate, $CaCO_3$, in an antacid tablet [ans: 0.50 g]
   (d) $1.2 \times 10^3$ mol of sodium hypochlorite, NaOCl [ans: $8.9 \times 10^4$ g]
   (e) 45 mmol of helium in a balloon [ans: $1.8 \times 10^{-1}$ g]

## 6.4 Summary

- The molar mass and average atomic mass of an element are numerically the same. This value is given on the periodic table.
- The unit of molar mass is g/mol, while the unit of average atomic mass is u.
- The molar mass of a compound is the sum of the molar masses of each entity in the compound.
- The amount, $n$, mass, $m$, and molar mass, $M$, of a pure substance are related to each other through the equation $n = \frac{m}{M}$.

## 6.4 Questions

1. Do 1.0 mol samples of different compounds all have the same mass? Explain your answer. K/U
2. Describe how chemists use mass as a way of getting a precise estimate of the number of entities in a sample. K/U
3. Why is an amount, obtained using the equation $n = \frac{m}{M}$, always an estimate rather than an exact value? K/U
4. Calculate the molar mass of each of the following compounds: T/I
   (a) iron(III) oxide, $Fe_2O_3$ (in rust)
   (b) calcium carbonate, $CaCO_3$ (in blackboard chalk)
   (c) octane, $C_8H_{18}$ (in gasoline)
   (d) calcium chlorate, $Ca(ClO_3)_2$ (in fireworks)
   (e) ammonium carbonate, $(NH_4)_2CO_3$ (in smelling salts)
5. Calcium chloride, $CaCl_2$, is used as a drying agent to protect electronics during shipping. Calculate the amount of calcium chloride in a 10.0 g sample. T/I
6. A typical energy drink contains 80.0 mg of the stimulant caffeine, $C_8H_{10}N_4O_2$. Calculate the amount of caffeine in the energy drink. T/I
7. A kernel of popping corn contains $1.22 \times 10^{-3}$ mol of water. What is the mass of this water? T/I
8. Calculate the amount of silicon in a 5.8 mg sample of pure silicon used to manufacture a computer chip. T/I
9. A tank truck carries 34 000 L of sulfuric acid. The density of sulfuric acid is 1.84 kg/L. T/I
   (a) What mass of sulfuric acid is in the truck?
   (b) What amount of sulfuric acid is in the truck?
10. A bag of intravenous fluid contains 0.154 mol of sodium chloride, NaCl. What mass of sodium chloride is required to prepare this bag? T/I
11. A 2.9 g sample of a pure compound contains 0.050 mol of the compound. Use this data to calculate the molar mass of the compound. T/I
12. A sodium atom has a mass of 22.99 u. Use Avogadro's constant and the conversion factor $1.661 \times 10^{-24}$ g/u to calculate the mass of 1.00 mol of sodium atoms. T/I