



## 4.2 | The Mole

---

Even the smallest amounts of ionic or molecular compounds used in the laboratory contain an enormous number of ions or molecules. Perhaps more than  $10^{21}$  molecules of water are present in a single drop of water! It is convenient (if not necessary) to have a special unit when counting atoms and molecules.

The chemist counting unit is the **mole**. A mole is defined as  $6.02 \times 10^{23}$  particles. The number  $6.02 \times 10^{23}$  is called **Avogadro's number**. The number of particles represented by Avogadro's number is called a mole, abbreviated **mol**.

Use Avogadro's number as a conversion factor when relating moles and particles.

$$1 \text{ mole} = 6.02 \times 10^{23} \text{ particles}$$

Two conversion factors:

$$\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ particles}} \quad \text{and} \quad \frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mol}}$$

### Worked Example 4-2

How many moles of sodium are in  $8.44 \times 10^{22}$  sodium atoms?

#### Solution

Given:  $8.44 \times 10^{22}$  Na atoms

Find: mol Na

Conversion Factor:

1 mol Na =  $6.02 \times 10^{23}$  Na (in this case "particles" are Na atoms)

What conversion factor should we apply?

$$\frac{1 \text{ mol Na}}{6.02 \times 10^{23} \text{ Na}} \quad \text{or} \quad \frac{6.02 \times 10^{23} \text{ Na}}{1 \text{ mol Na}}$$

We use 1 mol Na/ $6.02 \times 10^{23}$  atoms Na to cancel atoms Na.

$$8.44 \times 10^{22} \text{ Na atoms} \times \frac{1 \text{ mol Na}}{6.02 \times 10^{23} \text{ Na atoms}} = 0.140 \text{ mol Na}$$

### Practice 4-2

How many atoms are present in 0.0045 mol of xenon?

**Answer**

## 4.3 | Molar Mass

---

The **molar mass** of a substance is always numerically equal to the substance's formula mass with the unit of g/mol. For example, the formula mass of O<sub>2</sub> is 32.00 amu and the molar mass of O<sub>2</sub> is 32.00 g/mol (one mole of O<sub>2</sub> has a mass of 32.00 g).

### Worked Example 4-3

Calculate the formula mass and molar mass of each of the following:  
H<sub>2</sub>O, N<sub>2</sub>O<sub>3</sub>, and Mg(OH)<sub>2</sub>.

**Solution**

See Worked Example 4-1

$$\text{H}_2\text{O} = (2 \times 1.01 \text{ amu}) + (1 \times 16.00 \text{ amu}) = 18.02 \text{ amu}$$

$$\text{N}_2\text{O}_3 = (2 \times 14.01 \text{ amu}) + (3 \times 16.00 \text{ amu}) = 76.02 \text{ amu}$$

$$\text{Mg}(\text{OH})_2 = (1 \times 24.31 \text{ amu}) + (2 \times 1.01 \text{ amu}) + (2 \times 16.00 \text{ amu}) = 58.33 \text{ amu}$$

The **formula mass** of H<sub>2</sub>O is **18.02 amu** so the **molar mass** is **18.02 g/mol**.

The **formula mass** of N<sub>2</sub>O<sub>3</sub> is **76.02 amu** so the **molar mass** is **76.02 g/mol**.

The **formula mass** of Mg(OH)<sub>2</sub> is **58.33 amu** so the **molar mass** is **58.33 g/mol**.

**Practice 4-3**

Calculate the formula mass and molar mass of each of the following:  
 $\text{Fe}(\text{NO}_3)_2$  and  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ .

**Answer**

**Converting Between Grams and Moles**

We can use molar mass as a conversion factor to convert from grams to moles, and from moles to grams. For example, the molar mass of water is 18.02 g/mol. Thus:

$$1 \text{ mol H}_2\text{O} = 18.02 \text{ g H}_2\text{O}$$

From this equality we can write two conversion factors:

$$\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \quad \text{and} \quad \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}}$$

**Conversion of Grams to Moles**

Calculate the number of moles in 76.2 g of  $\text{H}_2\text{O}$ .

$$76.2 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 4.23 \text{ mol H}_2\text{O}$$

**Conversion of Moles to Grams**

Calculate the mass of 1.40 mol of  $\text{H}_2\text{O}$ .

$$1.40 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 25.2 \text{ g H}_2\text{O}$$

### Practice 4-4

Calculate the number of moles in 45.8 g of sucrose,  $C_{12}H_{22}O_{11}$ .

**Answer**

## 4.4 | Chemical Reactions and Chemical Equations

---

A **chemical reaction** is a process in which substances are transformed into new substances. Typical evidence of a chemical reaction is:

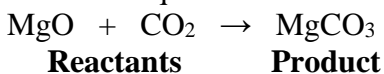
- A color change.
- A solid forms (precipitates).
- A gas forms (bubbles).
- Heat is given off or absorbed.

A **chemical equation** is a shorthand representation of a chemical reaction. In a chemical reaction **reactants** (starting materials) are converted into **products**.

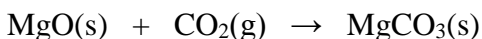
Consider the reaction in which magnesium oxide reacts with carbon dioxide to form magnesium carbonate.

We can represent the above “word description” by a “chemical equation”.

Chemical equation:



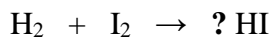
We often indicate the physical state of reactants and products using the following symbols: (s) for solid state; (l) for liquid state; (g) for gaseous state. If a substance is dissolved in water, it is an aqueous (aq) solution. States may or may not be given in chemical equations.



## 4.5 | Balancing Chemical Equations

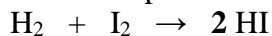
---

If you combine one molecule of  $\text{H}_2$  with one molecule of  $\text{I}_2$ , how many molecules of  $\text{HI}$  are produced?



A balanced equation contains the same number of each type of atom on each side of the equation.

A **coefficient** is a whole number placed in front of the formulas to balance the equation.



**2: coefficient**

**Use the following steps for balancing chemical equations**

Step 1: Count the atoms of each element on both sides of the equation.

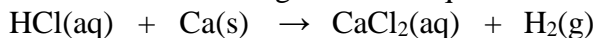
Step 2: Determine which elements are not balanced

Step 3: Balance one element at a time using coefficients.

Step 4: Do a final check, making sure ALL elements are balanced.

### Worked Example 4-4

Balance the following chemical equation:



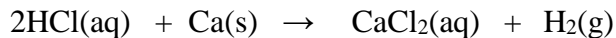
#### Solution

Step 1: Counting atoms

Reactants	Products
1 H atom	2 H atoms
1 Cl atom	2 Cl atoms
1 Ca atom	1 Ca atom

Step 2: Ca is balanced (one on each side). H and Cl are not balanced.

Step 3: Placing a 2 in front of  $\text{HCl}$  balances both the H and the Cl.



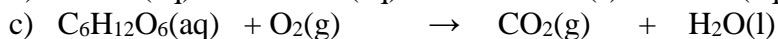
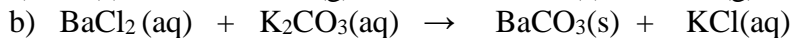
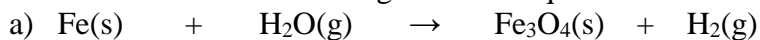
Step 4: ALL elements are balanced.

---

---

### Worked Example 4-5

Balance each of the following chemical equations:



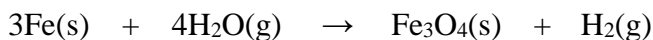
#### Solution

Step 1: Counting atoms

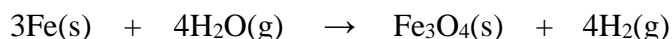
Reactants	Products
1 Fe atom	3 Fe atoms
2 H atom	2 H atoms
1 O atom	4 O atoms

Step 2: H is balanced (two on each side). Fe and O are not balanced.

Step 3: Placing a 3 in front of the Fe, balances the Fe. Placing a 4 in front of  $\text{H}_2\text{O}$  balances the O.



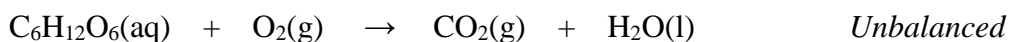
Step 4: H is not balanced now. Placing a 4 in front of  $\text{H}_2$  balances ALL elements.



b) We use the same procedure as in part (a).

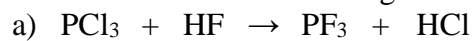


c) We use the same procedure as in part (a).



**Practice 4-5**

Balance each of the following chemical equations:



**Answer**



## 4.6 | Classification of Chemical Reactions

---

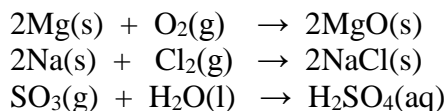
There is no comprehensive classification scheme that would accommodate all known chemical reactions. One approach is to classify reactions into four types: **combination**, **decomposition**, **single replacement** and **double replacement** reactions.

### I) Combination Reactions

In a combination reaction, two or more substances react to form a single product.

The general form of this reaction is  $(A + B \rightarrow AB)$ .

Some examples are shown below:

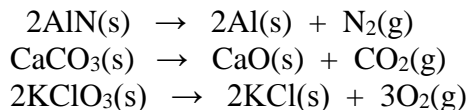


### II) Decomposition Reactions

In a decomposition reaction, a reactant splits into two or more simpler products.

The general form of the reaction is  $(AB \rightarrow A + B)$ .

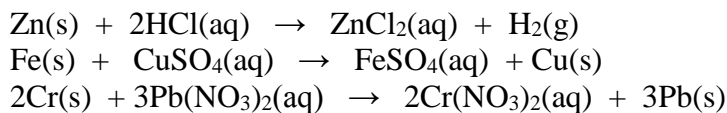
Some examples are shown below:



### III) Single Replacement Reactions

In a single replacement reaction, a reacting element switches place with an element in the other reacting compound. The general form of the reaction is  $(A + BC \rightarrow AC + B)$ .

Some examples are shown below:

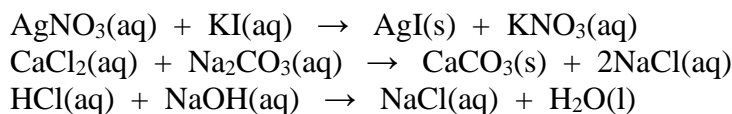


### IV) Double Replacement Reactions

In a double replacement reaction, two compounds exchange partners with each other to produce two different compounds.

The general form of the reaction is  $(AB + CD \rightarrow AD + BC)$ .

Some examples are shown below:



There are two common double-replacement reactions that we will study in this course:

1. Precipitation reactions
2. Acid-base reactions

We will study the precipitation reaction in the next section, and the acid-base reaction in *unit (6)*.

## 4.7 | Precipitation Reactions

---

A **precipitation reaction** involves the formation of an insoluble product, called a *precipitate*, when you mix solutions of two ionic substances. Most precipitations occur when the cations and anions of two ionic compounds change partners.

The first step toward understanding the precipitation reaction is to know which ionic compounds are soluble in water and which are not. In Table 4-1, some rules for the solubility of ionic compounds are listed. You should be able to use the solubility rules, but you will not be responsible for memorizing the table.

### The Solubility of Some Common Ionic Compounds (Table 4-1)

---

Soluble Compounds	Exception
<ul style="list-style-type: none"><li>• Sodium, potassium, and ammonium compounds</li><li>• Acetate and nitrates</li><li>• Hydrogen carbonates</li><li>• Chlorides, bromides, and iodides</li><li>• Sulfates</li></ul>	<p>Lead(II), silver, mercury(I) compounds</p> <p>CaSO<sub>4</sub>, BaSO<sub>4</sub>, PbSO<sub>4</sub></p>
Insoluble Compounds	Exception
<ul style="list-style-type: none"><li>• Carbonates and phosphates</li><li>• Hydroxides</li></ul>	<p>Sodium, potassium, and ammonium compounds</p> <p>NaOH, KOH, Ca(OH)<sub>2</sub>, and Ba(OH)<sub>2</sub></p>

### Worked Example 4-6

State whether each of the following compounds is soluble or insoluble in water.  
 $K_2SO_4$ ,  $Mg(NO_3)_2$ ,  $Pb(C_2H_3O_2)_2$ ,  $CaSO_4$ ,  $Ag_3PO_4$ ,  $Ba(OH)_2$ , and  $AgCl$ .

#### Solution

$K_2SO_4$	Soluble (contains potassium)
$Mg(NO_3)_2$	Soluble (contains nitrate)
$Pb(C_2H_3O_2)_2$	Soluble (contains acetate)
$CaSO_4$	Insoluble (contains sulfate “exception”)
$Ag_3PO_4$	Insoluble (contains phosphate)
$Ba(OH)_2$	Soluble (contains hydroxide “exception”)
$AgCl$	Insoluble (contains chloride “exception”)

### Practice 4-6

State whether each of the following compounds is soluble or insoluble in water.  
 $NaOH$ ,  $PbI_2$ ,  $Ba_3(PO_4)_2$ ,  $(NH_4)_2S$ ,  $CoCO_3$ ,  $Al(NO_3)_3$ ,  $Hg(OH)_2$ .

#### Answer

Let us consider two solid ionic compounds,  $AgNO_3$  and  $KCl$ . Both of these solids are soluble in water (see the solubility rules). Remember from *unit 3*, that ionic compounds are composed of ions, cations and anions.

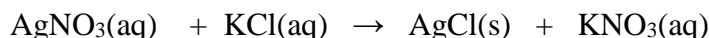
When  $AgNO_3$  is added to water, the ions separate and spread throughout the solvent. A  $AgNO_3$  solution, expressed as  $AgNO_3(aq)$ , does not contain any  $AgNO_3$  units, but rather silver ions ( $Ag^+$ ) and nitrate ions ( $NO_3^-$ ).

Similarly, when KCl is added to water, the ions separate and the KCl solution, expressed as KCl(aq), contains potassium ions (K<sup>+</sup>) and chloride ions (Cl<sup>-</sup>).

What happens when we mix these two solutions?

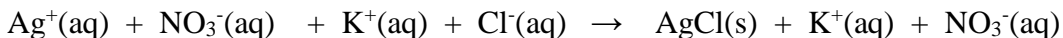
The instant that the solutions are mixed, all four ions (Ag<sup>+</sup>, NO<sub>3</sub><sup>-</sup>, K<sup>+</sup>, and Cl<sup>-</sup>) are present.

Assuming that cations change partners, the two possible products will be KNO<sub>3</sub> and AgCl. Next, we refer to the table to determine the solubilities of these two products. We find that KNO<sub>3</sub> is soluble and AgCl is insoluble. Applying the principles of equation writing, we can write:



The above equation, showing the complete chemical formulas, is called a **formula equation** because it shows the chemical formulas of the reactants and products without indicating their ionic characters.

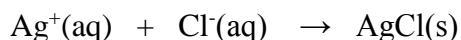
In describing reactions that occur in solutions it is desirable to write the equation for the reaction in ionic form, indicating explicitly the ions that actually exist in the solution:



The above equation with all soluble compounds shown as ions is called a **complete ionic equation**.

Notice that the K<sup>+</sup> and NO<sub>3</sub><sup>-</sup> ions don't undergo chemical changes. They are in the exact same form on both sides of the equation. Ions that don't undergo a chemical change during a chemical reaction are called **spectator ions**.

If we omit the spectator ions, we will have the **net ionic equation**:



Steps for Writing Formula Equation, Ionic Equation and Net Ionic Equation

1. Write the names, and then formulas of reactants (unit 3 naming).
2. Exchange cations and write the names, then formulas of products (unit 3 naming).
3. Write a chemical equation showing the formulas of reactants and products.
4. From the solubility rules include the (aq) for soluble and (s) for insoluble substances.
5. Balance the equation. This is a formula equation.
6. Write the substances with (aq) as ions. This is a total ionic equation.
7. Omit the spectator ions. This is a net ionic equation.

**Worked Example 4-7**

Write the formula equation, total ionic and net ionic equations for the reaction of  $\text{CaI}_2$  and  $\text{Na}_2\text{SO}_4$ .

**Solution**

1. Reactants: calcium iodide,  $\text{CaI}_2$ , and sodium sulfate,  $\text{Na}_2\text{SO}_4$ .
2. Products: calcium sulfate,  $\text{CaSO}_4$ , and sodium iodide,  $\text{NaI}$  (exchange cations).
3.  $\text{CaI}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + \text{NaI}$
4.  $\text{CaI}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + \text{NaI}(\text{aq})$
5.  $\text{CaI}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{NaI}(\text{aq})$
6.  $\text{Ca}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq}) + 2\text{Na}^{+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{Na}^{+}(\text{aq}) + 2\text{I}^{-}(\text{aq})$
7.  $\text{Ca}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{CaSO}_4(\text{s})$

**Practice 4-7**

Write the formula equation, total ionic and net ionic equations for the reaction of  $\text{MgCl}_2$  and  $\text{Na}_2\text{CO}_3$ .

**Answer**

## 4.8 | Redox and Nonredox Reactions

---

Reactions could also be classified as redox or nonredox reactions. The redox reactions are also called oxidation-reduction reactions.

**Redox reaction** is a reaction that involves electron transfer between two reacting substances.

**Nonredox reaction** is a reaction in which there is no electron transfer between reacting substances.

In an electron-transfer reaction, by definition, one substance must gain electrons while another substance must lose electrons.

**Oxidation** is the **loss** of electrons by a substance (alternatively, it is an *increase* in the *oxidation number* of a substance.)

**Reduction** is the **gain** of electrons by a substance (alternatively, it is a *decrease* in the *oxidation number* of a substance.)

We can classify the four reaction types (combination, decomposition, single replacement and double replacement reactions) as either redox or nonredox.

### Oxidation Number: Electron Bookkeeping

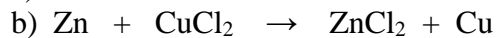
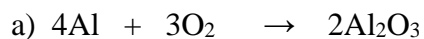
To determine whether electrons are transferred in a chemical reaction, we use a procedure that assigns an **oxidation number** to each atom in the reaction.

#### Some Shortcut Methods for Assigning Oxidation Number to Atoms

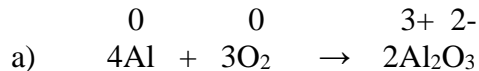
1. The oxidation number of hydrogen is usually +1. The oxidation number of oxygen is usually -2. The oxidation number of an element in its elemental state is zero.
2. In binary ionic compounds the oxidation number is predicted based on the position of metal and nonmetal in the periodic table. (Na = +1, Mg = +2, N = -3, F = -1).
3. In binary molecular compounds, the “more metallic” element (further down and/or to the left in the periodic table) tends to lose, and the “less metallic” tends to gain electrons.
4. The sum of the oxidation number of all atoms in a compound is zero; for a polyatomic ion, the sum is equal to the charge on the ion.

### Worked Example 4-8

In each of the following reactions, determine which element is oxidized and which is reduced:

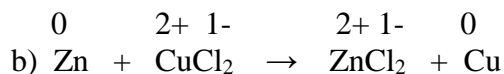


#### Solution



The oxidation number of aluminum changes from 0 to +3. This means that the aluminum has lost electrons, so the **aluminum is oxidized**.

The oxidation number of oxygen changes from 0 to -2. This means that the oxygen has gained electrons, so the **oxygen is reduced**.



The oxidation number of zinc changes from 0 to +2. This means that the zinc has lost electrons, so the **zinc is oxidized**.

The oxidation number of copper changes from +2 to 0 (it reduces). This means that the copper has gained electrons, so the **copper is reduced**.

The oxidation number of chlorine does not change, so it is neither oxidized nor reduced.

### Practice 4-8

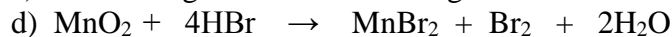
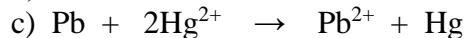
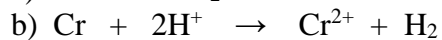
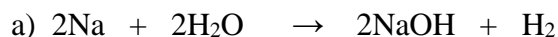
Assign oxidation number to the nitrogen in each of the following:



#### Answer

### Practice 4-9

In each of the following reactions, determine what is oxidized and what is reduced?



**Answer**

## 4.9 | Energy and Chemical Reactions

Besides changes in composition, energy changes always accompany chemical reactions. If the energy of reactants is higher than the energy of products, then heat energy is released. If the energy of reactants is lower than the energy of products, heat energy is absorbed.

Reactions are either exothermic or endothermic.

**Exothermic** reactions *release* heat energy and energy appears as a product.

**Endothermic** reactions *absorb* heat energy and energy is a reactant.

For example, the combustion of propane (a gas used as camping fuel) is an exothermic reaction. Heat (specifically, 531 kcal per mole of propane) is released as a product when  $\text{CO}_2$  and  $\text{H}_2\text{O}$  are formed. See the chemical equation below, note the energy on the product side.



The reaction of lead(II) oxide with carbon is an endothermic reaction. In this example, heat, 26 kcal per mole of lead(II) oxide, is required and shown as a reactant:



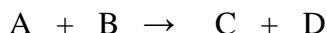


## 4.10 | Reversible Reactions and Equilibrium

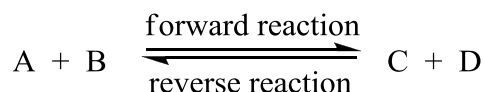
---

Theoretically, all chemical reactions can proceed in both the forward and reverse directions. However, we consider many reactions as irreversible and assumed that reactants are completely converted to products.

A single arrow, pointing in one direction, is used to indicate an irreversible reaction.



Two arrows pointing in opposite directions (a double arrow notation) are used to indicate a reversible reaction.



A **reversible reaction** proceeds in both the forward and a reverse direction. The forward reaction is called “the reaction to the right”, and the reverse reaction is called “the reaction to the left.”

Let’s examine a typical reversible reaction using the equation above. When we add compound A to compound B, we initiate the forward reaction. The two compounds begin to react at a certain rate (the rate of the forward reaction). As the reaction proceeds, the concentrations of reactants (A and B) decrease while the concentration of products (C and D) increase.

Now, compounds C and D begin to react and initiate the reverse reaction at a certain rate (the rate of reverse reaction).

Since we have a large amount of reactants at the beginning, the rate of the forward reaction is the highest at the start of the reaction and then gradually decreases.

Conversely, the rate of the reverse reaction, which begins at zero, gradually increases. At some point, the rate of the forward and reverse reactions will become equal. At this point the concentration of reactants and products remain **constant** and a state of **equilibrium** has been reached. We refer to this as **dynamic equilibrium** because the reactions continue in both directions without producing an overall change in the concentrations of reactants and products.

## 4.11 | Mass Relationship in Chemical Equations

---

The mass relationships in chemical reactions are called **stoichiometry** (*stoy-key-ah-meh-tree*).

Stoichiometry is derived from the Greek words *stoicheion* meaning “element” and *metron* meaning “measure”.

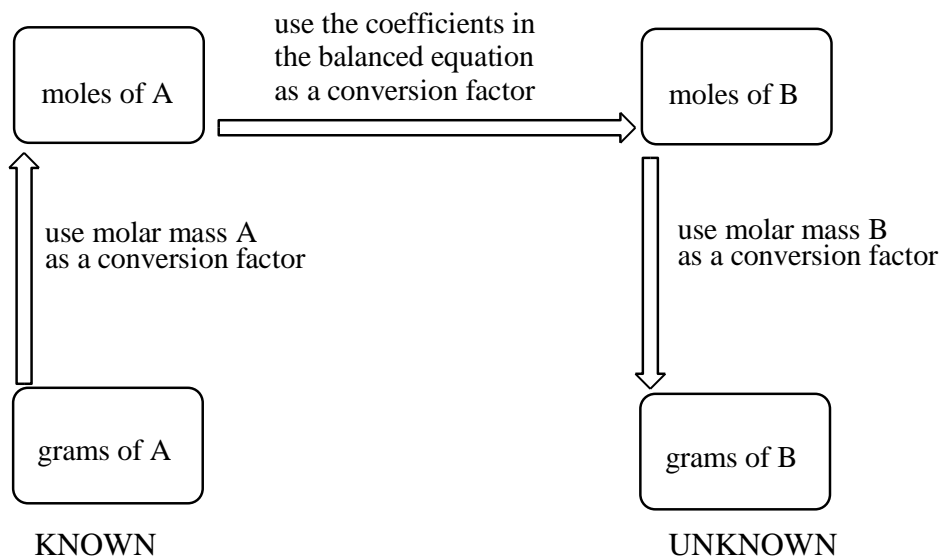
In a typical stoichiometric problem, the mass of one substance in the reaction is given (known) and the mass of another substance in the reaction is asked (unknown). This is called a **mass-mass problem**.

After balancing the chemical equation, we solve the mass-mass stoichiometry problems as follows:

- 1) Convert the mass of the known to moles of known using the molar mass of the known as a conversion factor (periodic table).
- 2) Convert the moles of known to the moles of unknown using the coefficients in the balanced equation.
- 3) Convert the moles of the unknown to the mass using the molar mass of the unknown.

Suppose that the mass of A is given (known) and the mass of B is asked (unknown).

For a general reaction ( $A \rightarrow B$ ), the flowchart outline for this type of calculation is:



### Worked Example 4-9

What mass of H<sub>2</sub>O would be produced from 7.4 g of O<sub>2</sub>?

Given the equation:  $2\text{H}_2(\text{g}) + 1\text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

#### Solution

- a) First convert the given mass of O<sub>2</sub> to moles of O<sub>2</sub> using the molar mass of O<sub>2</sub> as a conversion factor, (molar mass of O<sub>2</sub> = 2 x 16.00 g/mol = 32.00 g/mol).

$$7.4 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.23 \text{ mol O}_2$$

- b) Next, convert the moles of O<sub>2</sub> to moles of H<sub>2</sub>O by applying a mole ratio using the coefficient in the balanced equation, (1 mol O<sub>2</sub>.... 2 mol H<sub>2</sub>O).

$$0.23 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} = 0.46 \text{ mol H}_2\text{O}$$

- c) Lastly, use the molar mass of H<sub>2</sub>O (18.02 g/mol) as a conversion factor to calculate the mass of H<sub>2</sub>O produced.

$$0.46 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 8.3 \text{ g H}_2\text{O}$$

After you gain confidence in solving stoichiometry problems, you should be able to perform a continuous calculation (dimensional analysis):

grams O<sub>2</sub> → mol O<sub>2</sub> → mol H<sub>2</sub>O → grams H<sub>2</sub>O

$$7.4 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 8.3 \text{ g H}_2\text{O}$$

### Practice 4-10

How many grams of Cl<sub>2</sub> can be produced from 3.1 g of HCl?

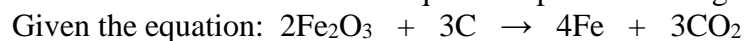
Given the equation:  $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + 2\text{H}_2\text{O} + 1\text{Cl}_2$

Perform a continuous calculation.

#### Answer

**Practice 4-11**

Calculate the mass of carbon required to produce 18.6 g of iron.

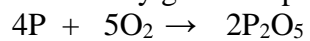


Perform a continuous calculation.

**Answer**

**Practice 4-12**

How many grams of phosphorus are required to react completely with 6.6 g O<sub>2</sub>?



Perform a continuous calculation.

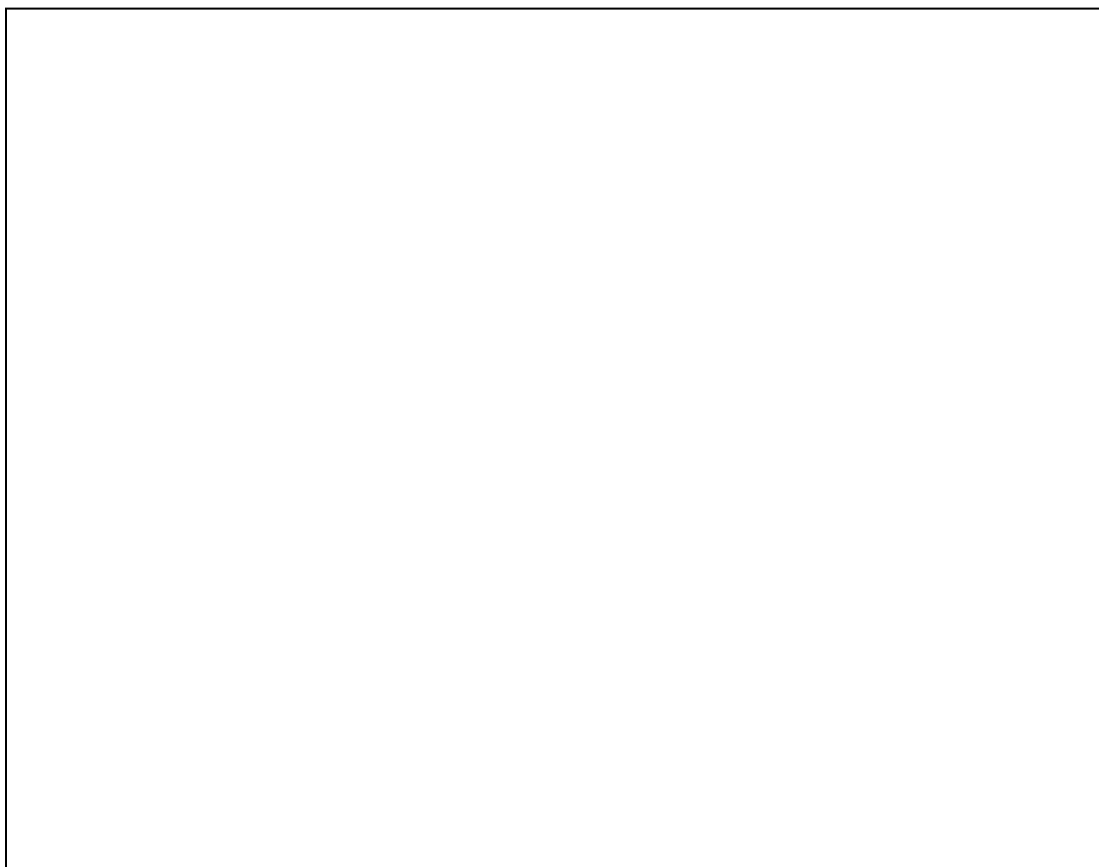
**Answer**

**Practice 4-13**

Given the equation:  $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$

Calculate the following:

- How many moles of Fe are produced from 1.80 mol of CO?
- How many moles of  $\text{Fe}_2\text{O}_3$  are needed to produce 66.0 g of  $\text{CO}_2$ ?
- How many grams of Fe will be produced from 7.52 g of CO?

**Answer**

## 4.12 | Gases

The **Ideal Gas Law**:

$$PV = nRT$$

P = Pressure (units of atm, torr, or mm Hg) 1 atm = 760 mm Hg

V = Volume (units of L or mL)

n = number moles of gas

T = Temperature in Kelvin (K)

R = Universal Gas Constant = 0.08206 L·atm/(mol·K)

The ideal gas law predicts how atoms or molecules in a gaseous state will behave at the microscopic level. All gases obey the law at very low pressures.

### Worked Example 4-10

An O<sub>2</sub> cylinder contains 2.00 x 10<sup>2</sup> kg of liquid oxygen. Calculate the number of liters of oxygen at 20.0°C and 1.00 atm of pressure.

#### Solution

a) Ideal Gas Law: PV = nRT, rearrange and solve for volume: V = nRT/P

b) Convert kg of O<sub>2</sub> into moles O<sub>2</sub> using the molar mass of O<sub>2</sub> and 1 kg = 1000 g as conversion factors (molar mass of O<sub>2</sub> = 2 x 16.00 g/mol = 32.00 g/mol).

$$2.00 \times 10^2 \text{ kg O}_2 \times \frac{1000 \text{ g O}_2}{1 \text{ kg O}_2} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 6.25 \times 10^3 \text{ mol O}_2$$

c) Convert °C into K using:

$$K = ^\circ\text{C} + 273.15$$

$$K = 20.0^\circ\text{C} + 273.15$$

$$K = 293.15 \text{ K}$$

d) Substitute universal gas constant (R), and values from b and c, into equation derived in a) and solve for V:

$$V = nRT/P$$

$$V = \frac{(6.25 \times 10^3 \text{ mol O}_2)(0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(293.15 \text{ K})}{1.00 \text{ atm}}$$

$$V = 1.50 \times 10^5 \text{ L}$$

**Boyle's Law:** At constant temperature, volume is inversely proportional to pressure.  
 $V \sim 1/P$  therefore  $PV = \text{constant}$

**Charles's Law:** At constant pressure, volume is directly proportional to temperature.  
 $V \sim T$  therefore  $V/T = \text{constant}$

**Avogadro's Law:** At constant P and T, volume is directly proportional to # of moles  
 $V \sim n$  therefore  $V/n = \text{constant}$

Solving Gas Law Problems:

Boyle's Law             $P_1V_1 = P_2V_2$   
 Charles's Law         $V_1/T_1 = V_2/T_2$   
 Combined Gas Law    $P_1V_1/T_1 = P_2V_2/T_2$   
 Avogadro's Law       $V_1/n_1 = V_2/n_2$

At Standard Temperature and Pressure (STP,  $T = 0^\circ\text{C}$  and  $P = 1 \text{ atm}$ ), 1 mole of a gas occupies 22.4 L

Calculate the missing quantity:

$P_1 = 0.600 \text{ atm}$	$V_1 = 205 \text{ mL}$	$P_2 = ?$	$V_2 = 246 \text{ mL}$
$V_1 = 750. \text{ mL}$	$T_1 = 200. \text{ K}$	$V_2 = 1000. \text{ mL}$	$T_2 = ?$
$V_1 = x \text{ mL}$	$n_1 = 5.0 \text{ mol}$	$V_2 = ?$	$n_2 = 10 \text{ mol}$

- Use Boyle's Law:  $P_2 = 0.500 \text{ atm}$
- Use Charles's Law:  $T_2 = 267 \text{ K}$
- Use Avogadro's Law:  $V_2 = 2x \text{ mL}$

A gas at  $20.^\circ\text{C}$  and  $1.10 \text{ atm}$  fills a  $4.5 \text{ L}$  container. Calculate the volume this gas will occupy at  $120.^\circ\text{C}$  and  $2.20 \text{ atm}$ .

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(1.10 \text{ atm})(4.5 \text{ L})}{293 \text{ K}} = \frac{(2.20 \text{ atm})(V_2)}{393 \text{ K}}$$

$$V_2 = \frac{(1.10 \text{ atm})(4.5 \text{ L})(393 \text{ K})}{(293 \text{ K})(2.20 \text{ atm})}$$

$$V_2 = 3.0 \text{ L}$$

A gas at 20. °C exerts a pressure of 0.50 atm and occupies a volume of 96.0 L Calculate the number of moles of the gas.

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(0.50 \text{ atm})(96.0 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K})(293 \text{ K})}$$

$$n = 2.0 \text{ mol}$$

## Homework Problems

---

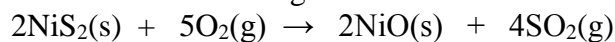
- 4.1 Calculate the number of moles in a 5.00-g sample of each of the following:
- Caffeine ( $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$ )
  - Aspirin ( $\text{C}_9\text{H}_8\text{O}_4$ )
  - PenicillinV ( $\text{C}_{16}\text{H}_{18}\text{N}_2\text{O}_5\text{S}$ )
- 4.2 Balance the following equations:
- $\text{Cl}_2\text{O}_7 + \text{H}_2\text{O} \rightarrow \text{HClO}_4$
  - $\text{Fe}_2(\text{CO}_3)_3 \rightarrow \text{Fe}_2\text{O}_3 + \text{CO}_2$
  - $\text{H}_3\text{PO}_4 + \text{Mn}(\text{OH})_2 \rightarrow \text{Mn}_3(\text{PO}_4)_2 + \text{H}_2\text{O}$
  - $\text{P}_2\text{H}_4 \rightarrow \text{PH}_3 + \text{P}_4$
  - $\text{SO}_2\text{Cl}_2 + \text{HI} \rightarrow \text{H}_2\text{S} + \text{H}_2\text{O} + \text{HCl} + \text{I}_2$
- 4.3 Write the formula equation, total ionic and net ionic equations for each of the following:
- The reaction of  $\text{Pb}(\text{NO}_3)_2$  and  $\text{LiCl}$ .
  - The reaction of  $\text{ZnSO}_4$  and  $\text{KOH}$ .
- 4.4 In the following redox reactions, identify what is oxidized and what is reduced:
- $\text{Zn}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$
  - $\text{Sn}^{2+}(\text{aq}) + 2\text{Ce}^{4+}(\text{aq}) \rightarrow \text{Sn}^{4+}(\text{aq}) + 2\text{Ce}^{3+}(\text{aq})$
  - $2\text{Au}(\text{s}) + 6\text{H}^+(\text{aq}) \rightarrow 2\text{Au}^{3+}(\text{aq}) + 3\text{H}_2(\text{g})$
  - $4\text{Co}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Co}_2\text{O}_3(\text{s})$
  - $2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$



4.5 Define, explain, or describe each of the following terms:

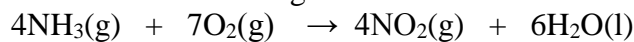
- a. exothermic reaction
- b. endothermic reaction
- c. reversible equation
- d. dynamic equilibrium

4.6 Consider the following reaction:



- a. How many moles of  $\text{SO}_2$  will be produced from 8.95 g of  $\text{O}_2$ ?
- b. How many grams of  $\text{NiO}$  would be formed from the reaction of 0.0125 mol of  $\text{NiS}_2$ ?
- c. How many grams of  $\text{O}_2$  are needed to completely react with 3.74 g of  $\text{NiS}_2$ ?

4.7 Consider the following reaction:



- a. How many moles of  $\text{NH}_3$  react with 5.64 mol of  $\text{O}_2$ ?
- b. How many moles of  $\text{NO}_2$  are obtained from 3.27 mol of  $\text{O}_2$ ?
- c. How many moles of  $\text{H}_2\text{O}$  will be produced from 8.95 g of  $\text{NH}_3$ ?
- d. How many grams of  $\text{NH}_3$  will be needed to produce 0.0160 g of  $\text{NO}_2$ ?

4.8 Calculate the volume occupied by 1.5 moles of an ideal gas at  $25^\circ\text{C}$  and a pressure of 0.80 atm. ( $R = 0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K})$ ).

4.9 A sample of carbon monoxide has a volume of 150 mL at  $10.^\circ\text{C}$  and 0.75 atm. What pressure will be exerted by the gas if the temperature is increased to  $522^\circ\text{C}$  and the volume changed to 500. mL?