Chapter 12

WHY ARE GASES SO IMPORTANT?

Our atmosphere is vital to life on Earth. Without an atmosphere, we could not exist. There are also other gases, some beneficial and some toxic, which we encounter in our lives. In addition, we will see that a simple mathematical formula enables the conversion of pressure, volume, and temperature of a gas to moles.

12.1 Background

We live in a gaseous world. We rely on the atmosphere to provide a livable climate and oxygen for our lungs. Yet our atmosphere does not extend very far from the surface of the Earth. Figure 12.1 gives a view of the thinness of our atmosphere.

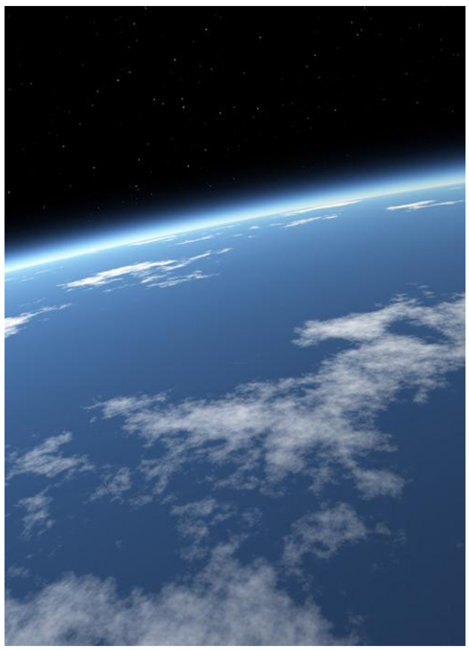

Figure 12.1 A partial view of the Earth from space showing the thinness of the atmosphere.

The chemical composition of our atmosphere is completely different to that of all the other planets as is shown in Table 12.1. For the inner (rocky) planets, it is typical for the atmosphere to be mostly carbon dioxide, while for the outer (gaseous) planets, hydrogen and helium predominate with some methane (CH_4). The Earth's atmosphere is unique in having high levels of oxygen gas which makes Earth a livable planet. The Earth's atmosphere also has a very low concentration of carbon dioxide – about 0.04% which means that the so-called greenhouse effect is currently only enough to help keep the planet warm but not too hot. For comparison, at the surface of Venus, the concentration of carbon dioxide is so high that the surface temperature is close to 500ºC.

Table 12.1 Atmospheric composition of some of the planets in our solar system

Planet	Atmospheric Composition (percent)		
Venus	96.5% CO_2	3.5% N_2	
Earth	78.1% N_2	20.9% O_2	0.9% Ar
Mars	95.3% CO_2	2.7% N_2	
Jupiter	86.2% H_2	13.6% He	0.2% CH_4
Saturn	96.3% H_2	3.3% He	0.4% CH_4

12.2 Common Properties of Gases

Though each gas has its unique features, all gases share certain properties and these provide the topic of this section. The properties relate back to Chapter 2, Section 2.3, on Phases of Matter while their explanation is in terms of Section 2.4, the Kinetic-Molecular Model of Matter.

Gases Have Very Low Densities

When a substance is in the gas phase, the molecules have overcome the inter-molecular attractions and are moving freely. At ordinary pressures, most of a gas-filled container is empty space. As a result, the density of a gas is less than one-thousandth of the density of a liquid.

Gases Mix in Any Proportion

Gas molecules are free to move, so placing different gases together will result in the gases mixing. The mixing process can be slow. Thus it is possible to collect a gas denser than air, such as carbon dioxide, in a beaker and study its properties. Similarly, it is possible to collect a gas less dense than air, such as helium, provided the beaker is inverted. Unless the mouth of the beaker is covered in each case, the collected gas will mix with the surrounding air in a few minutes.

Volume of a Gas Changes with Pressure

Whereas solids and liquids are almost completely incompressible, gases can be compressed as they consist of mostly empty space. If the volume of the container is reduced (at constant temperature), then the gas molecules will have a shorter distance to travel before they impact the walls of the container again. As pressure is the number of impacts per second on the container walls, the pressure goes up as the volume goes down. Conversely, the pressure goes down as the volume increases and the molecules have farther to travel between impacts.

Pressure of a Gas Changes with Temperature

Temperature is a measure of the average kinetic energy of the gas molecules. Thus as the temperature goes up (at constant volume, such as in a spray can), the average kinetic energy ($\frac{1}{2}mv^2$) will increase. As the mass is constant, it is the molecular velocities that will increase. If the molecules are moving faster, they will impact the container walls more frequently, resulting in an increase in pressure.

Volume of a Gas Changes with Temperature

If the pressure is held constant, when the temperature of a gas is increased, then the volume will be affected. As the molecules increase in velocity, then the volume of the gas must increase, so the molecules have to travel correspondingly further between collisions with the container walls.

12.3 Gases That Are Elements

In Chapter 10, among the chemical elements that were described, eleven of them are gases. Some properties of the gaseous chemical elements are shown in Table 12.2.

Table 12.2 Gaseous chemical elements and a summary of their properties.

	Element	Properties
Group 0	Hydrogen, H_2	Colourless, odourless, flammable, low density.
Group 15	Nitrogen, N_2	Colourless, odourless, unreactive.
Group 16	Oxygen, O_2	Colourless, odourless, supports combustion.
	Ozone, O_3	Colourless, metallic-smell, poisonous.
Group 17	Fluorine, F_2	Yellow, pungent, highly reactive, poisonous.
	Chlorine, Cl_2	Green, pungent, reactive, poisonous.
Group 18	Helium, He	Colourless, odourless, unreactive.
	Neon, Ne	Colourless, odourless, unreactive.
	Argon, Ar	Colourless, odourless, unreactive.
	Krypton, Kr	Colourless, odourless, unreactive.
	Xenon, Xe	Colourless, odourless, unreactive.
	Radon, Rn	Colourless, odourless, radioactive, unreactive.

12.4 Gases That Are Oxides

In addition to some gaseous elements, there are many non-metal oxides that are gases. The most common ones are discussed below.

Carbon Monoxide

Carbon monoxide is a colourless, odourless, gas which is highly poisonous. It is haemoglobin molecules which pick up oxygen molecules from the lungs and transport them to where they are needed in the body. However, carbon monoxide bonds more strongly to haemoglobin than does oxygen and, as a result, the carbon monoxide blocks the pickup of oxygen. The haemoglobin can no longer function as a conveyer of oxygen and the individual loses consciousness and dies. Thus only a low concentration of carbon monoxide will kill.

When carbon-containing compounds burn, one of the products is carbon monoxide. The proportion of carbon monoxide is higher when there is insufficient air to form carbon dioxide. Wood-burning stoves and cigarettes produce carbon monoxide, as do poorly-serviced oil, propane, and natural gas, heating systems. It is for this reason that, in most jurisdictions, carbon monoxide detectors are mandatory in any home or dwelling using combustion for heating or cooking.

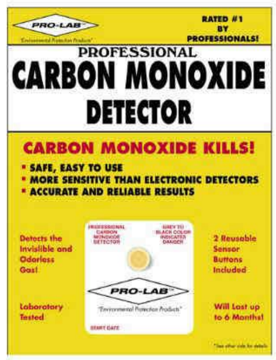

Figure 12.2 A carbon monoxide detector is essential for any home which uses combustion for heating or cooking

Carbon monoxide is flammable, burning to form carbon dioxide in a combination reaction:

$$2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$$

Carbon Dioxide

Like carbon monoxide, carbon dioxide is colourless and odourless. However, it is not poisonous. Carbon dioxide does not burn, nor does it support combustion. Plants absorb carbon dioxide in photosynthesis, while carbon dioxide is the gas that we exhale from our metabolic processes. The gas is significantly denser than air, though as mentioned above, over time all gases mix.

In Chapter 9, Section 9.7, Double Replacement Reactions, it was shown how carbon dioxide can be produced by reacting a metal carbonate with a dilute acid. The double replacement reaction producing carbonic acid is followed by the decomposition of carbonic acid to form carbon dioxide and water. This reaction is illustrated here by the commonly-used addition of dilute hydrochloric acid to solid calcium carbonate:

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2CO_3(aq)$$

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

Carbon dioxide is an infra-red radiation-trapping gas, or as is commonly called, a "greenhouse gas." It is through the presence of carbon dioxide, water vapour (another radiation-trapping gas), and others that our planet has a livable temperature. The problem, as everyone is aware, is that with all the combustion processes of the last 150 years, the concentration of carbon dioxide has been increasing rapidly (Figure 12.3). It is accepted by qualified scientists that global climate change is occurring, the only matter for debate is the rate at which this is happening.

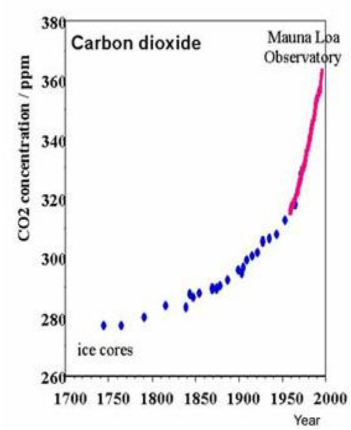

Figure 12.3 The change in atmospheric concentration of carbon dioxide with time.

As carbon dioxide dissolves in water in a combination reaction to form carbonic acid:

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

there is concern that, with rising levels of carbon dioxide in the atmosphere, more will dissolve in the seas, making them more acidic, possibly to the point at which coral reefs will start to dissolve.

Dinitrogen Oxide

Nitrogen forms many oxides, but three in particular are important. Dinitrogen oxide, also called nitrous oxide, is colourless but it is unique in having a sweet-smelling odour. It is also the only common gas apart from oxygen that supports combustion. For this reason, dinitrogen oxide is used in some types of high performance race cars and motorcycles to be mixed with the fuel in the engine in place of air.

Dinitrogen oxide is used as an anesthetic gas, such as for tooth extractions and other simple medical procedures. Inhalation causes temporary unconsciousness, though in smaller concentrations it causes involuntary giggles, hence its common name of laughing gas.

Nitrogen Monoxide

Until recent times, nitrogen monoxide, a colourless gas, was regarded as an unimportant pollutant. Now we know that inside our bodies, this gas plays a crucial role of transmitting nerve signals to the brain. Nitrogen monoxide is also used in hospitals to treat 'blue baby syndrome.' The medical term is infant pulmonary hypertension, which means that an inadequate quantity of blood flows through the lungs and not enough oxygen is absorbed, causing some babies to turn blue.

Dinitrogen oxide is very unusual in that opening a container to the air, it reacts with oxygen to form brown clouds of nitrogen dioxide:

$$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$$

Nitrogen Dioxide

Nitrogen dioxide is one of the very few coloured gases, being deep brown. It has a very sharp, acidic odour and is poisonous. Nitrogen dioxide is a major contributor to urban pollution.

Sulfur Dioxide

Sulfur dioxide enters the atmosphere from volcanoes and from industrial processes, such as the burning of coal and sulfur-rich oils. Sulfur dioxide is a colourless gas with a sharp, acidic smell/taste. It dissolves in water to give acidic sulfurous acid:

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

Summary of Common Gaseous Oxides

The properties of common gaseous oxides are summarized in Table 12.3.

Table 12.3 Some gaseous oxides and a summary of their properties.

	Oxide	Properties
Group 14	Carbon monoxide, CO	Colourless, odourless, flammable, poisonous.
	Carbon dioxide, CO_2	Colourless, odourless, slightly acidic.
Group 15	Dinitrogen oxide, N_2O	Colourless, sweet-smelling, supports combustion.
	Nitrogen monoxide, NO	Colourless, reacts with oxygen.
	Nitrogen dioxide, NO_2	Red-brown, sharp acidic odour, acidic, poisonous.
Group 16	Sulfur dioxide, SO_2	Colourless, 'acidic' smell, acidic, poisonous.

12.5 Gases That Are Hydrides

There are five common gases that are non-metal hydrides.

Methane

Methane, CH_4, is colourless, odourless, and very flammable. It is the major component of natural gas. Methane, like carbon dioxide, is an infra-red radiation-trapping gas. Thus increases in methane, too, (Figure 12.4) are contributing to climate change. There are several sources of the increased methane. The paddy fields used for rice production are one significant source as anaerobic bacteria thrive in the water-covered soil producing methane as a waste product. Likewise, the bacteria in the stomachs of ruminant animals, such as sheep and cows, generate large quantities of methane.

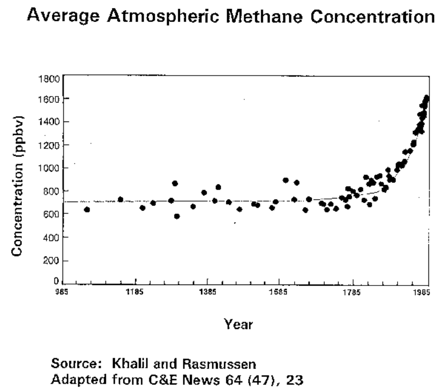

Figure 12.4 A plot of the increase in atmospheric methane with time

AMMONIA

Ammonia, NH_3, is a colourless, pungent-smelling gas. It is the only common gas that is basic. Ammonia has this property because it dissolves in water to form a solution of ammonium hydroxide.

$$NH_3(aq) + H_2O(l) \rightarrow NH_4OH(aq)$$

As a result of being basic, the solution is a good household cleaner. Ammonia is used in large quantities in agricultural areas to improve crop yields (Figure 12.5).

Figure 12.5 Ammonia is pumped into the soil to improve crop yield.

HYDROGEN SULFIDE

Hydrogen sulfide, H_2S, is colourless with a very strong 'rotten-egg' smell. The gas is poisonous, in very low concentrations causing headaches and in higher concentrations, death. It is formed by bacteria in anaerobic environments, such as bogs and swamps. Underground deposits of natural gas often have very high concentrations of hydrogen sulfide, and the natural gas is then referred to as 'sour gas.'

HYDROGEN CHLORIDE

Hydrogen chloride, HCl, is a colourless, sharp-smelling gas which dissolves in water to form acidic hydrochloric acid:

$$HCl(g) \xrightarrow{\text{dissolve in water}} HCl(aq)$$

HYDROGEN CYANIDE

Hydrogen cyanide, HCN, is a colourless gas with a faint 'almond' odour. It is very poisonous. One source of low concentrations of hydrogen cyanide is tobacco smoke (Figure 12.6).

Figure 12.6 Tobacco smoke contains measurable concentrations of poisonous hydrogen cyanide.

SUMMARY OF COMMON GASEOUS HYDRIDES

In Table 12.4, some properties of the common gaseous hydrides are summarized:

Table 12.4 Some gaseous hydrides and a summary of their properties.

	Hydrides	Properties
Group 14	Methane, CH_4	Colourless, odourless, flammable.
Group 15	Ammonia, NH_3	Colourless, pungent odour, basic.
Group 16	Hydrogen sulfide, H_2S	Colourless, 'rotten-egg' smell, very poisonous.
Group 17	Hydrogen chloride, HCl	Colourless, sharp odour, very acidic.
Other	Hydrogen cyanide, HCN	Colourless, faint 'almond' odour, very poisonous.

12.6 Units of Pressure

Pressure is a scientific measure in many aspects of our lives. Measuring our blood pressure is an important component of our health care. Checking the air pressure in vehicle tires can prevent an accident. When going up in an aircraft, the cabin has to be pressurized so that there is a high-enough concentration of air for us to survive.

THE SI UNIT OF PRESSURE

Pressure is a derived unit (see Chapter 3, Section 3.9, Volume) – it is the force per unit area, the greater the force on a surface (such as our skin), the higher the pressure. The scientific unit for pressure measurement is the Pascal, named after the French scientist, Blaise Pascal (Figure 12.7).

A pressure of one Pascal is very small, so pressures are usually measured in kilopascals (kPa). Meteorologists sometimes measure pressure in the related units of millbars where 1000 millibars = 100 kPa.

Figure 12.7 The S.I. unit of pressure is named after the French scientist, Blaise Pascal.

The standard for pressure is now 100.0 kPa but the previous reference value was the standard atmosphere. However, as atmospheric pressure is always fluctuating, there really is no 'standard' atmosphere. The chosen value was the pressure needed to support a column of mercury that is 760 mm high, sometimes called 760 torr after the Italian scientist Torricelli. The modern equivalent pressure to one standard atmosphere is 101.325 kPa.

STANDARD CONDITIONS

When reporting scientific measurements, it is common to cite them at a specific combination of temperature and pressure, so that readings made by different scientists and laboratories can be compared. Such accepted values of pressure and temperature are known as ***standard conditions***. All measurements are now reported at ***Standard Ambient Temperature and Pressure*** (SATP) which is 100.0 kPa pressure and 298 K (25ºC) temperature. Values were previously reported at Standard Temperature and Pressure which is 101.325 kPa pressure and 273 K (0ºC) temperature.

12.7 Vapour Pressure

In Chapter 2, Section 2.3, Phases of Matter, the term 'evaporation' was used for the conversion of a liquid to the gas phase below the boiling point and 'sublimation' for the conversion from the solid phase to the gas phase. Both of these phenomena can be explained in terms of moving particles, see Chapter 2, Section 2.4, the Kinetic-Molecular Model of Matter. As the molecules escape into the gas phase, they will exert a pressure called the ***vapour pressure***. The higher the temperature, the more molecules will go from the liquid (or solid) phase to the gas phase. That is, the vapour pressure of a substance increases with increasing temperature. The substance will begin to boil when its vapour pressure equals or exceeds the atmospheric pressure. Thus the boiling point of a substance is not fixed, but it depends upon the atmospheric pressure at that time. The vapour pressure curve for water is shown in Figure 12.8.

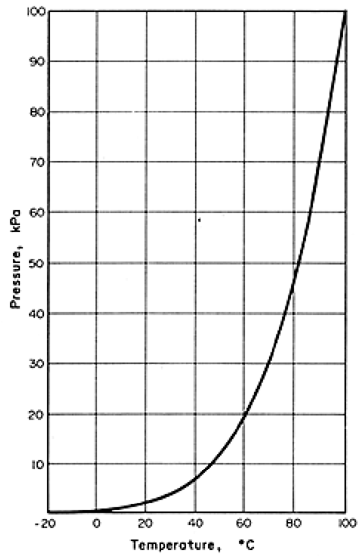

Figure 12.8 The vapour pressure curve for water.

Three sample values for the boiling point of water are shown in Table 12.5. Thus, at sea level, the boiling point of water is usually slightly less than 100ºC and only equals that value when the atmospheric pressure is about 101.4 kPa.

Table 12.5 The dependence of boiling point of water upon atmospheric pressure

Boiling point	Pressure	Comment
99.97ºC	101.325 kPa	Former standard pressure
99.61ºC	100.000 kPa	Current standard pressure
approx. 69 ºC	approx. 26 kPa	Top of Mt. Everest

12.8 The Ideal Gas Equation

The three things that we can measure about a gas are its volume, temperature, and pressure. There is a simple formula which relates these three values to the number of moles of gas:

$$PV = nRT$$

where:

P = pressure in kiloPascals (kPa)
V = volume in litres (L)
n = number of moles (mol)
R = the Universal Gas Constant (8.31 $kPa \cdot L \cdot mol^{-1} \cdot K^{-1}$)
T = temperature in Kelvin (K)

Knowing this formula, and the relationship between mass and moles, it is possible to calculate any missing numerical value as can be seen in the following example.

EXAMPLE 12.1

3.25 g of nitrogen gas are placed in a 1.50 L container at a temperature of 23ºC. What is the pressure inside the container?

Answer

<u>Strategy</u>	<u>Relationship</u>
Mass of N_2 → mol N_2	1 mol $N_2 \equiv 28.0$ g
Mol/vol/temp of N_2 → pressure N_2	P = nRT/V

$$\text{Mol } N_2 = 3.25 \text{ g } N_2 \times \left(\frac{1 \text{ mol}}{28.0 \text{ g}}\right) = 0.116 \text{ mol } N_2$$

$$\text{Pressure } N_2 = \frac{0.116 \text{ mol} \times 8.31 \text{ kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 296 \text{ K}}{1.50 \text{ L}} = 190 \text{ kPa}$$

IDENTIFYING AN ELEMENT

The Ideal Gas Equation can also be used to identify an element as shown in the next calculation.

EXAMPLE 12.2

When 1.09 g of a gas, EF_4, where E is an unidentified element, is placed in a container of volume 125 mL at a temperature of 24ºC, it exerts a pressure of 104 kPa. Calculate the molar mass of the unknown gas and identify the element E.

Answer

<u>Strategy</u>	<u>Relationship</u>
Press/vol/temp of EF_4 → mol EF_4	n = PV/RT
Mass/mol of EF_4 → molar mass EF_4	m.m. = m/n

$$\text{Mol } EF_4 = \frac{104 \text{ kPa} \times 0.125 \text{ L}}{8.31 \text{ kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 297 \text{ K}} = 5.26 \times 10^{-3} \text{ mol}$$

$$\text{molar mass } EF_4 = \frac{1.09 \text{ g}}{5.26 \times 10^{-3} \text{ mol}} = 207 \text{ g} \cdot \text{mol}^{-1}$$

$$\text{molar mass E} = (207 - 4 \times 19.0) \text{ g} \cdot \text{mol}^{-1} = 131 \text{ g} \cdot \text{mol}^{-1}$$

Element E must be xenon, Xe.

MOLECULAR FORMULA OF A COMPOUND

In Chapter 8, Section 8.5, Empirical and Molecular Formulas, it was shown how to calculate an empirical formula, then, given the molar mass, calculate the molecular formula. The process is shown in Figure 12.9.

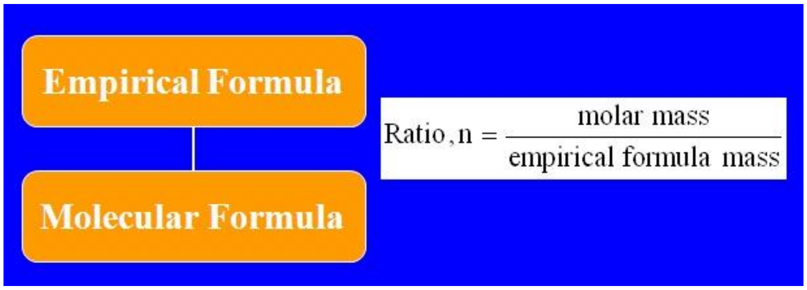

Figure 12.9 To find the molecular formula from the empirical formula, the evaluation of a mass ratio is needed.

Using the Ideal Gas Equation, given the mass, volume, temperature, and pressure, the molar mass of a gas can be calculated and thus the molecular formula derived. This process is shown in the next example.

EXAMPLE 12.3

A compound consists of 49.0% carbon, 2.7% hydrogen, and 48.3% chlorine. When 9.87 g of the compound is vapourized in a 1.50 L container, the gas exerts a pressure of 148 kPa at a temperature of 125ºC. Calculate the empirical formula and the molecular formula of the compound.

Answer

Assume 100.0 g of compound. This will contain 49.0 g of carbon, 2.7 g of hydrogen, and 48.3 g of chlorine.

$$\text{Mol of C} = 49.0\ \text{g C} \times \left(\frac{1\ \text{mol}}{12.0\ \text{g}}\right) = 4.08\ \text{mol C}$$

$$\text{Mol of H} = 2.7\ \text{g H} \times \left(\frac{1\ \text{mol}}{1.01\ \text{g}}\right) = 2.7\ \text{mol H}$$

$$\text{Mol of Cl} = 48.3\ \text{g Cl} \times \left(\frac{1\ \text{mol}}{35.5\ \text{g}}\right) = 1.36\ \text{mol Cl}$$

$$\text{Ratio} = \frac{4.08\ \text{mol C}}{1.36\ \text{mol}} : \frac{2.7\ \text{mol H}}{1.36\ \text{mol}} : \frac{1.36\ \text{mol Cl}}{1.36\ \text{mol}} = 3.00\ \text{C} : 2.0\ \text{H} : 1\ \text{Cl}$$

The empirical formula is C_3H_2Cl and the corresponding empirical formula mass is 73.5 g·mol^{-1}. The molecular formula will be some multiple of this: $(C_3H_2Cl)_n$. To find *n*, the ratio of the molar mass and the empirical formula mass is needed. The molar mass can be calculated using the Ideal Gas Equation:

Strategy	Relationship
Press/vol/temp of gas → mol gas	n = PV/RT
Mass/mol of gas → molar mass gas	m.m. = m/n

$$\text{Mol gas} = \frac{148\ \text{kPa} \times 1.50\ \text{L}}{8.31\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 398\ \text{K}} = 6.71 \times 10^{-2}\ \text{mol}$$

$$\text{molar mass gas} = \frac{9.87\ \text{g}}{6.71 \times 10^{-2}\ \text{mol}} = 147\ \text{g} \cdot \text{mol}^{-1}$$

$$\text{Mass ratio}, n = \frac{147\ \text{g}}{73.5\ \text{g}} = 2$$

The molecular formula is $(C_3H_2Cl)_2$ or more correctly, $C_3H_4Cl_2$

DENSITY OF A GAS

In Chapter 3, Section 3.10, Density, it was shown how to determine the density of a solid or liquid. Using the Ideal Gas Equation, it is possible to calculate the density of any gas at a specific combination of pressure and temperature. To do this, the density relationship is used in terms of molar mass and molar volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\text{molar mass}}{\text{molar volume}}$$

EXAMPLE 12.4

Calculate the density of sulfur dioxide gas at a pressure of 101 kPa and a temperature of 22ºC

Answer

$$\text{Density} = \frac{\text{molar mass}}{\text{molar volume}}$$

Molar mass SO_2 = 64.1 g

$$\text{Molar volume, V} = \frac{\text{nRT}}{\text{P}} = \frac{1\ \text{mol} \times 8.31\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 295\ \text{K}}{101\ \text{kPa}} = 24.3\ \text{L}$$

$$\text{Density} = \frac{64.1\ \text{g}}{24.3\ \text{L}} = 2.64\ \text{g} \cdot \text{L}^{-1}$$

12.9 Where Next?

Gases are important in our lives, and so are solutions, the topic of Chapter 13.