CHAPTER 5
PHYSICAL BEHAVIOR OF MATTER

In studying chemistry we study the composition of matter, and how it changes. Energy changes accompany these changes in matter.

ENERGY

Although there are many more sophisticated definitions, **energy** is probably best defined as the capacity to do work. **Work** is done whenever a force is used to change the position and/or motion of matter in a system.

All energy can be classified into two types—*potential energy* and *kinetic energy*. **Potential energy** is stored energy. **Kinetic energy** is energy of motion.

When we roll a ball up a hill, we expend kinetic energy. The ball, however, acquires potential energy as a result of the work done on it. At the top of the hill, the ball at rest has potential energy due to its position. If the ball rolls down the hill, the potential energy will change to the kinetic energy of motion.

Forms of Energy

Energy, whether potential or kinetic, is associated with physical phenomena. Traditionally, different *forms* of energy are identified as follows:

- **Thermal energy** is the total amount of internal energy an object has due to the random motion of its molecules.
- **Chemical energy** is the energy associated with chemical change involving the making and breaking of bonds.
- **Light energy** is the energy associated with electromagnetic radiation (visible, ultraviolet, infrared, x-rays, etc.)
- **Nuclear energy** is the energy associated with changes in the masses of atoms' nuclei.
- **Heat energy** is a transfer of energy (usually thermal) from a body of higher temperature to a body of lower temperature.
- **Electrical energy** and **mechanical energy** are forms less important in chemical reactions.

Conservation of Energy

During ordinary chemical or physical changes, energy is conserved. Energy may be changed from one form to another, or transferred from one body or system to another, but the total amount of energy remains constant.

Any activity that releases energy is said to be **exothermic.** Activities that require energy are said to be **endothermic**.

Measurement of Energy

Temperature is a measure of the *average* kinetic energy of the particles in a substance. Temperature is not a form of energy. Two things should be emphasized in this definition:

1. The word *average* indicates that temperature might be calculated by adding the kinetic energies of all the particles and dividing by the number of particles. Under most conditions, each particle possesses its own kinetic energy, which may be quite different from that of any of its neighboring particles.

2. The word *heat* does *not* appear in the definition, although the temperature may be affected by the amount of heat in the system. A rising temperature is associated with the addition of heat to the system; falling temperature is associated with a loss of heat. Whenever two bodies with different temperatures get close enough to each other, *heat* flows from the body with the higher temperature to the one with the lower temperature until the two bodies are at the same temperature.

A **calorimeter** is a device used to measure heat of reaction. The calorimeter uses the change in temperature of water to measure the amount of heat produced during exothermic processes and the amount of heat absorbed during endothermic processes. The process is allowed to proceed in a container that is in contact with a known mass of water. The temperature of the water is measured before and after the reaction (process), and the change in temperature is determined.

Temperature change must be converted to units of heat energy. A traditional unit of heat energy is the **calorie**. One calorie is defined as the quantity of heat needed to raise the temperature of 1 gram of water by 1 Celsius degree. Using standard units, the amount of heat required to raise 1 gram of water by 1 Celsius degree is 4.18 **Joules**.

A conversion constant known as the *specific heat* can make the same equation work for either unit and for substances other than water. For water, the specific heat is either 1.00 calorie/gram°C or 4.18 joule/g.°C.

The equation is: $q = m\,C\,\Delta T$ where:

q = heat gained or lost
m = mass
C = specific heat
ΔT = temperature change

SAMPLE PROBLEMS

1. How many Joules are absorbed by 30.0g of water when its temperature is raised from 20.°C to 40.°C?

Solution:

Given: mass (water) = 30.0g
$\Delta T = 40.°C - 20.°C = 20.°C$

Known: C = 4.18 J/g. °C (for water)

Find: heat gained

Equation: Heat gained $= mC\Delta T$
$= 30.0g \times 4.18\ J/g°C \times 20.°C$
$= \mathbf{2.5 \times 10^3\ J}$

2. A reaction chamber inside a calorimeter contains 150.g of water at 19°C. A reaction is permitted to take place in the chamber. After the reaction, the temperature of the water is 29°C. How many calories were released by the reaction?

Heat released = mass × specific heat × temperature change
$= mC\Delta T$
$= 150.g \times 1\ cal/g°C \times 10.°C$
$= \mathbf{1.5 \times 10^3}$ **calories**

QUESTIONS

Answer the following questions using Tables B, C, and D of the *Reference Tables for Physical Setting/Chemistry.*

1. The temperature of a substance is a measure of its particles'

(1) average potential energy (2) average kinetic energy
(3) enthalpy (4) entropy

2. How many kiloJoules (kJ) are equivalent to 1.0 Joule?

(1) 0.001 kJ (2) 0.01 kJ (3) 1000 kJ (4) 10,000 kJ

3. What is the total number of Joules of heat energy absorbed when the temperature of 200. grams of water is raised from 10°C to 40°C?

(1) 125 J (2) 836 J (3) 25000 J (4) 33400 J

4. How many grams of water will absorb a total of 5.02×10^3 Joules of energy when the temperature of the water changes from 10.0°C to 30.0°C?

(1) 10.0 g (2) 20.0 g (3) 30.0 g (4) 60.0 g

5. How many Joules of heat energy are absorbed in raising the temperature of 10. grams of water from 5.0°C to 20.°C?

(1) 6.3×10^2 (2) 2.5×10^2 (3) 1.2×10^2 (4) 3.0×10^1

6. The temperature of 50 grams of water was raised to 50°C by the addition of 4200 Joules of heat energy. What was the initial temperature of the water?

(1) 10°C (2) 20°C (3) 30°C (4) 60°C

MATTER

The traditional definition of **matter** is anything that has mass and volume. The term matter is ordinarily used when referring to any material found in nature. These materials are found in what seems to be an infinite variety of forms, and are classified in many ways. Two groups into which all matter may be placed are *homogeneous matter* and *heterogeneous matter.*

- **Homogeneous matter** includes all those materials having uniform characteristics throughout a given sample. This group includes elements, compounds, and solutions.
- **Heterogeneous matter** is not uniform in its composition. Its parts are not alike.

Scientists also classify matter into the two groups—*pure substances* and *mixtures* of substances. **Pure Substances** are almost always homogeneous. **Mixtures** may be homogeneous or heterogeneous. Solutions are examples of homogeneous mixtures.

Pure Substances

Pure substances can be elements or compounds. Elements are composed of atoms, all of which have the same atomic number. They cannot be broken down into other substances by chemical means. Examples are O_2 and Ne. Compounds are made of two or more elements but have a definite composition of elements. Examples of compounds are H_2O and CO_2. Pure substances have constant composition and constant properties which can be used in their identification. (Examples of such properties are: melting point, boiling point, solubility, and chemical reactivity.)

PHASES OF MATTER

Matter exists in nature in three **phases**—solid, liquid, and gas. These terms might be used to describe materials as they are found in nature under ordinary conditions. Water is most often found in liquid form, air is a mixture of gases, and the rocky portion of the earth consists mostly of materials in solid form.

In the consideration of the phases of matter, we shall be concerned with the transfer of heat as different materials undergo *phase changes*—that is, changes from one phase to another. A change in temperature or pressure can make a pure substance change from one phase to another without changing its chemical identity. The kinetic molecular theory of gases can be used not only to model gases, but it can also explain the properties of solids and liquids.

Kinetic Molecular Theory of Gases. Since so many gases exhibit the same behavior, a model was developed to help explain this similarity. This kinetic molecular theory of gases is based on the following assumptions:

1. All gases are composed of tiny, individual particles called molecules that are in continuous motion. These particles move rapidly, randomly and in straight lines.
2. When particles collide with one another, energy is transferred *without loss,* from one particle to the other. Therefore, the net total energy of the system remains constant.
3. Compared to the distances between them, the particles are so small that their volumes are considered to be zero.
4. The particles have no attraction for one another.

Gases that have these properties are known as **ideal gases.**

PROPERTIES OF GASES, LIQUIDS AND SOLIDS

Gases

Gases are transparent, they can be compressed and they expand without limit. Gases assume the shape and volume of their container. The volume of a given mass of a gas is the volume of the container that holds it.

Gases exert pressure on the walls of a container. Pressure is most often expressed in units of force per unit area of surface.

Earth's atmosphere exerts pressure, called air pressure, or atmospheric pressure. Air pressure is measured with a barometer. One type of barometer consists of a glass tube about one meter long, sealed at one end and partly filled with mercury. The open end of the tube is placed in a reservoir of mercury that is exposed to the atmosphere. The weight of the air exerts a pressure on the surface of the mercury in the reservoir. This pressure supports a column of mercury in the glass tube.

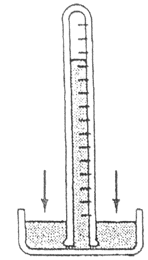


Figure 5-1. Mercury Barometer. The greater the air pressure (arrows) the higher the mercury rises in the tube.

The height of the mercury column is directly related to the amount of pressure exerted by the atmosphere on the mercury in the reservoir. At 0°C at sea level, the atmospheric pressure supports a column of mercury 760 mm high. In honor of E. Torricelli, who invented the mercury barometer, this pressure is often called 760 torr. This value—760 torr or 760 mm of mercury—is also known as **one atmosphere (1.00 atm)** of pressure, or **standard pressure (STP)**. It is equal to **101.3 kilopascals (kPa)**.

The kinetic molecular theory of gases explains pressure as the force of collision of the moving gas molecules against the walls of the container. Gases are greatly compressible because of the empty spaces which exist between the molecules. Gases assume their container's shape because the molecules of gases are not connected.

Liquids

Liquids have definite volumes but do not have definite shapes. Like gases, liquids take on the shape of their containers. Liquids, however, have low compressibility.

The kinetic molecular theory explains these properties. Liquids have low compressibility because the distances between molecules are not large. The molecules of liquids are packed close together, virtually touching each other. Liquids do not have their own shapes because the molecules are not rigidly connected to each other.

Evaporation. The force of attraction among particles of a liquid is much greater than among particles of a gas. The particles in a liquid are in constant motion, their speed depending on temperature. Particles that have a greater-than-average speed can overcome forces of attraction and leave the surface of the liquid to enter the gas, or vapor, phase. The term **vapor** is used to refer to the gas phase of a substance that is usually a solid or liquid at room temperature.

The escape of particles from the surface of a liquid to form the vapor is called **evaporation.** In an open container, evaporation continues until all the liquid has vaporized.

Condensation. If a liquid is placed in a closed container at a given temperature, only a small quantity will evaporate. It is impossible for particles that escape into the gas phase to leave the container. As more and more particles enter the gas phase, the likelihood that some of them will strike the liquid surface and return to the liquid phase increases. The change from vapor to liquid is called **condensation.** In a closed system, the rate of condensation eventually becomes equal to the rate of evaporation at the given temperature.

Vapor Pressure. In any closed system containing a liquid, the vapor produced by evaporation exerts a pressure, which is called the **vapor pressure** of the liquid. Every liquid has its own characteristic vapor pressure. Vapor pressure increases as temperature increases.

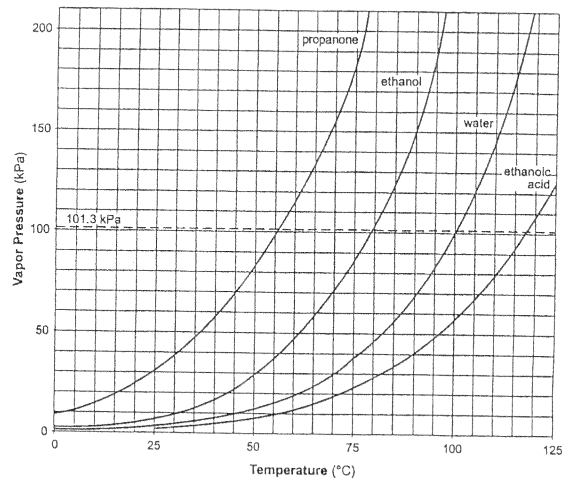


Figure 5-2. Vapor Pressure Curves for Four Liquids.

Boiling Point. Upon being heated, most liquids eventually boil. Boiling is characterized by the formation of bubbles of vapor in the body of the liquid, as the change from liquid to vapor takes place within the body of the liquid, as well as at the surface.

As described earlier, vapor pressure increases with an increase in temperature. The temperature at which the vapor pressure of a liquid is equal to the pressure pushing down on the surface of the liquid is called the **boiling point** of the liquid. The **normal boiling point** of a liquid is defined as the temperature at which the vapor pressure of the liquid is one atmosphere, or 760 torr or 101.3 kPa.

In most circumstances, when reference is made to the boiling point of a substance, it is the normal boiling point that is indicated. However, the temperature at which a liquid will boil can be changed by changing the pressure pushing down on the surface of the liquid. For example, the vapor pressure of water at 80°C is 47 kPa. Thus, if the pressure on the surface of a sample of water is reduced to 47 kPa, the water will boil at 80°C. Conversely, increasing the pressure raises the boiling point temperature.

Heat of Vaporization and Heat of Condensation. The particles that escape from the surface of a liquid are those having the highest kinetic energy. The remaining particles have lower average kinetic energy. Thus, heat must be added to the liquid to maintain constant temperature. Vaporization, then, is an endothermic process.

$$\text{heat} + H_2O(\ell) \rightarrow H_2O(g)$$

The **heat of vaporization** of a liquid is the number of calories per gram of liquid that must be added in order to maintain a constant temperature while vaporization, or boiling, occurs. For water at its normal boiling point of 100°C, the heat of vaporization is 539 calories per gram, or 2259 Joules per gram.

Heat of condensation is the number of calories per gram that must be *removed* in order to maintain constant temperature during condensation. Steam condensing on your hand produces severe burns due to the heat liberated on condensation. Steam at 100°C is much more damaging than water at the same temperature. The process of condensation is exothermic.

$$H_2O(g) \rightarrow H_2O(\ell) + \text{heat}$$

For water, the heat of condensation is 539 calories per gram or 2259 Joules per gram – the same as the heat of vaporization.

Solids

Solids have a definite volume and a definite shape. The particles in a solid have an orderly arrangement in regular geometric patterns, called crystals. Some apparent solids (referred to as amorphous solids - super cooled liquids) do not have regular geometric patterns. Some examples are: glass, plastics, and wax. Although the particles in a solid do not move freely about throughout the solid, they are in constant motion. They vibrate around their fixed positions. The molecules of solids are close together, similar to liquids, so they are only slightly more compressible than liquids.

Freezing-Melting Point. When a liquid is cooled (heat is removed), a temperature is eventually reached at which the liquid begins to freeze. It changes to a solid. This temperature, which remains constant until all the liquid has solidified at 1 atmosphere pressure, is called the **freezing point** of the liquid. While the liquid is cooling, the average kinetic energy of its particles decreases until it is low enough for the attractive forces to be able to hold the particles in the fixed positions characteristic of the solid phase.

Alternately, warming a solid eventually causes the solid to melt. It begins to change to a liquid. During this phase change, the temperature remains constant until all the solid has liquefied at 1 atmosphere pressure. This temperature is called the **melting point** of the solid. For any given substance, the melting point temperature is exactly the same as the freezing point temperature. The only difference is in the direction of approach.

The melting-freezing point of a substance may also be defined as the temperature at

Heat of Fusion and Heat of Solidification. Melting is an endothermic process. In order to maintain a constant temperature during this phase change, heat must be continually added to the system.

$$heat + H_2O(s) \rightarrow H_2O(\ell)$$

The amount of heat needed to change a unit mass of a substance from solid to liquid at constant temperature and 1 atmosphere pressure is called the **heat of fusion** of that substance. The heat of fusion of ice at 0°C and 1 atmosphere is 334 Joules per gram. To change one gram of ice at 0°C and 1 atmosphere pressure to one gram of water at the same conditions, 334 Joules of heat must be added. Conversely, to change 1 gram of liquid water at 0°C and 1 atm to 1 gram of ice at the same conditions, 334 Joules of heat must be removed. This is the **heat of solidification**.

$$H_2O(\ell) \rightarrow H_2O(s) + heat$$

Because of the principle of the conservation of energy, the heat of solidification of a substance is equal to the heat of fusion. If 334 J/g. must be added to melt one gram of ice, we must remove 334 J/g to refreeze it.

A graph of temperature vs time will make it easier to visualize the phase change process. Figure 5-3 is such a graph for a substance with a melting point of 50°C and a boiling point of 110°C. The portion of the curve between A and B represents the time the solid phase is being heated. Note the uniform increase in temperature during this interval. The portion of the curve between B and C represents the time the substance is undergoing a phase change, from solid to liquid (heat of fusion). Note that there is no change in temperature during this interval, even though heat is being added at a constant rate. The heat is being absorbed in the process of the phase change.

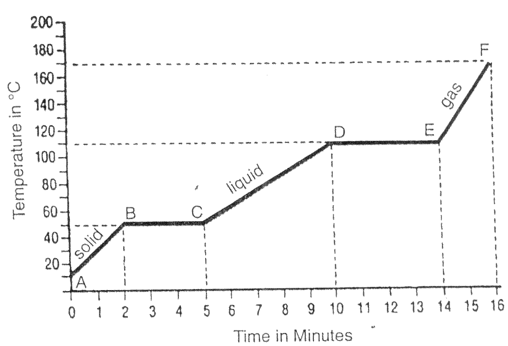


Figure 5-3. Heating Curve

The part of the curve between C and D represents the time interval during which the liquid is being heated. Note the uniform increase in temperature during this interval. At point D, the temperature stops rising.

The interval between D and E represents the time interval in which the liquid is changing to a gas (heat of vaporization). Once again, there is no change in temperature during this time. At point E, all of the liquid has changed to a gas, and the temperature of the gas begins to rise at a uniform rate.

Such a graph is called a *heating curve*. If you were to read this graph from right to left, it would be a *cooling curve*.

Sublimation. Under certain conditions, it is possible for a substance to change from a solid directly into the gas phase without obviously passing through the liquid phase. This solid-to-gas change is called **sublimation.** Iodine and carbon dioxide are examples of substances that may undergo sublimation. Iodine and carbon dioxide have small non-polar molecules with weak intermolecular attractions.

QUESTIONS

Answer the following questions using Tables A, B, and H of the *Reference Tables for Physical Setting/Chemistry.*

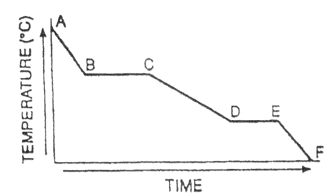
1. The heat of vaporization for water at its normal boiling point is

(1) 0.069 kJ/g (2) 0.25 kJ/g (3) 10.6 kJ/g (4) 2.26 kJ/g

2. Which conditions of pressure and temperature exist when ice melts at its normal melting point?

(1) 1 atm and 0°C (2) 760 atm and 0°C
(3) 1 atm and 0 K (4) 760 atm and 273 K

3. The graph below represents the uniform cooling of water at 1 atmosphere, starting with water as a gas above its boiling point.



Which segments of the cooling curve represent the freezing temperature?

(1) AB and CD (2) DE (3) AB and EF (4) CD and EF

4. The heat of fusion of a substance is the energy measured during a

(1) phase change
(2) temperature change
(3) chemical change
(4) pressure change

5. Which material exists as a supercooled liquid at STP?

(1) salt (2) sand (3) diamond (4) glass

6. Which substance will sublime at room temperature (20°C)?

(1) $C_{12}H_{22}O_{11}(s)$
(2) $C_6H_{12}O_6(s)$
(3) $SiO_2(s)$
(4) $CO_2(s)$

7. When the temperature of a sample of water is changed from 20°C to 100°C, the change in its vapor pressure is

(1) 2.7 kPa
(2) 93 kPa
(3) 98 kPa
(4) 101.3 kPa

8. As ice at 0°C changes to water at 0°C, the average kinetic energy of the ice molecules

(1) decreases (2) increases (3) remains the same

9. Which process occurs when dry ice, $CO_2(s)$, is changed into $CO_2(g)$?

(1) crystallization
(2) condensation
(3) sublimation
(4) solidification

10. As water in a sealed container is cooled from 20°C to 10°C, its vapor pressure

(1) decreases (2) increases (3) remains the same

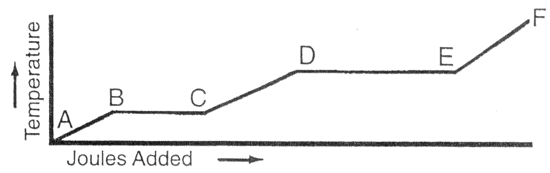
11. Water will boil at a temperature of 40°C when the pressure on its surface is

(1) 1.9 kPa (2) 2.2 kPa (3) 7.3 kPa (4) 101.3 kPa

12. A 1-gram sample of which substance in a sealed 1-liter container will occupy the container completely and uniformly?

(1) Ag(s) (2) Hg(ℓ) (3) $H_2O(\ell)$ (4) $H_2O(g)$

Base your answer to questions 13 and 14 on the graph below, which represents the uniform heating of a water sample at standard pressure, starting at a temperature below 0°C.



13. The number of Joules required to vaporize the entire sample of water at its boiling point is represented by the interval between

(1) A and B (2) E and F (3) C and D (4) D and E

14. If 5.0 grams of water undergoes a temperature change from C to D, the total energy absorbed is

(1) 340 Joules (2) 420 Joules (3) 760 Joules (4) 2100 Joules

15. The vapor pressure of ethanol at its normal boiling point is

(1) 10.7 kPa (2) 13.3 kPa (3) 36.4 kPa (4) 101.3 kPa

16. Which substance *cannot* be decomposed by a chemical change?

(1) mercury (II) oxide
(2) potassium chlorate
(3) water
(4) copper

17. Which change of phase represents fusion?

(1) gas to liquid
(2) gas to solid
(3) solid to liquid
(4) liquid to gas

18. Which is the equivalent of 750. Joules?

(1) 0.750 kJ (2) 7.50 kJ (3) 75.0 kJ (4) 750. kJ

IDEAL GAS LAWS

Ideal gases are those whose gas particles (molecules):

1. travel in random, constant, straight lines.

2. are separated by great distances relative to their sizes; the volume of the gas particles is considered negligible.

3. have no attractive forces between them.

4. have collisions that may result in a transfer of energy between particles, but the total energy of the system remains constant.

The properties of gases can be readily explained with the kinetic molecular theory. The thermal energy of gas molecules is due to their random motion. The temperature is their average kinetic energy which depends on both their mass and speed. The more mass or speed, the greater the average kinetic energy of the molecules (or temperature).

If a constant amount of a gas like oxygen is heated, its mass cannot change so the molecules of oxygen must go faster to be at a higher temperature. If they go faster, they should collide harder and more frequently on the sides of the container. Therefore the higher the **temperature**, the higher the **pressure** for a constant volume.

If we want the pressure to stay constant when the temperature increases, the volume must become larger so that faster moving molecules are farther apart. Even though the molecules at higher temperature are moving faster, their lower concentration compensates so the pressure can stay constant. As a result, as the **temperature** of a gas increases, for it to remain at constant pressure, its **volume** must increase.

At constant temperature if the volume of a gas is decreased, the concentration of molecules increases making more frequent collisions. As a result, as the **volume** decreases the **pressure** increases.

At 0°C (the freezing point of water) the molecules of water are still moving. They can go slower, so 0°C cannot be **absolute zero**, where there is no kinetic energy of the molecules. Absolute zero is 0 Kelvin, or 0 K. To change to the Kelvin (absolute) temperature scale 273° must be added to the Celsius temperature

$$K = °C + 273°$$

The freezing point of water is 273 K and the boiling point is 373 K at one atmosphere. In all gas law problems the temperature must be in Kelvin degrees.

The properties can be combined into one equation known as the **combined gas law** for an ideal gas.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Note:
If temperature is constant $T_1 = T_2$, the equation becomes $P_1 V_1 = P_2 V_2$
If the pressure is constant $P_1 = P_2$, the equation becomes

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

If the volume is constant $V_1 = V_2$ and the equation becomes

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

The advantages of using this equation are:

- Only one equation is needed
- All initial values are on the left of the equation; all new values are on the right of the equation.
- In cases where either temperature or pressure is unchanged, it can be ignored (or included); its value will not affect the problem.

It is also recommended that tables be set up and used. The following model can be used:

	1 **Initial conditions**	2 **New or final conditions**
Volume		
Pressure		
Temperature K		

SAMPLE PROBLEM

A sample of gas had a volume of 200. L under a pressure of 0.300 atm and a temperature of 20.° C. What volume would the sample occupy at a pressure of 0.250 atm and a temperature of 30.°C?

Solution:

1. Convert temperatures to Kelvin:

20.°C + 273 = 293K; 30.°C + 273 = 303K

2. Set up a table.

	1	2
V	200. L	x
P	0.300 atm	0.250 atm
T	293 K	303K

3. Write equation: $\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2}$ or $V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1}$

4. Substitute:

$$V_2 = 200.\text{ L} \times \frac{0.300\text{ atm}}{0.250\text{ atm}} \times \frac{303\text{K}}{293\text{K}}$$

= 248 L

Avogadro's Hypothesis. Two equal volumes of gases at the same temperature and pressure must have equal numbers of particles, even though their masses are not the same. For example, at the same temperature and pressure, one liter of nitrogen gas contains the same number of molecules as one liter of hydrogen gas, even though the mass of the nitrogen gas is fourteen times that of the hydrogen gas.

Since the two gases are at the same temperatures, they cannot have the same average speeds. Since the heavier nitrogen molecules exert the same pressure as the lighter hydrogen, the nitrogen must be moving at a slower speed. At the same temperature, lighter molecules travel faster than heavier molecules.

Deviations from the Gas Laws. Real gases do not behave exactly as predicted by the "ideal" gas of the model. Deviations from the ideal behavior are due to the fact that gas particles *do have volume* and they *do exert some attraction* for one another. Deviations from the gas laws are least obvious among light gases at high temperatures and low pressures. These conditions are optimum for high kinetic energies and maximum separation of particles. Hydrogen and helium are closest to "ideal" gases.

Standard Temperature and Pressure. Gas volumes are influenced considerably by changes in temperature and pressure. Thus, when working with gases, it is convenient to define standard reference conditions. By convention, 0°C (273K) and 760 torr or 101.3 kPa (1 atm) represent **standard temperature** and **pressure**, often designated as **STP**. When dealing with a gas volume, STP is assumed unless otherwise indicated.

QUESTIONS

Answer the following questions using Tables A, B, C, and D of the *Reference Tables for Physical Setting/Chemistry.*

1. The boiling point of water at standard pressure is

(1) 0.000 K (2) 100. K (3) 273 K (4) 373 K

2. Which temperature represents absolute zero?

(1) 0 K (2) 0°C (3) 273 K (4) 273°C

3. A 100.-milliliter sample of helium gas is placed in a sealed container of fixed volume. As the temperature of the confined gas increases from 10.°C to 30.°C, the internal pressure

(1) decreases (2) increases (3) remains the same

4. If the pressure on 36.0 milliliters of a gas at STP is changed to 0.190 atm at constant temperature, the new volume of the gas is

(1) 9.00 mL (2) 126 mL (3) 189 mL (4) 226 mL

5. At constant temperature the pressure on 8.0 liters of a gas is increased from 1 atmosphere to 4 atmospheres. What will be the new volume of the gas?

(1) 1.0 L (2) 2.0 L (3) 32 L (4) 4.0 L

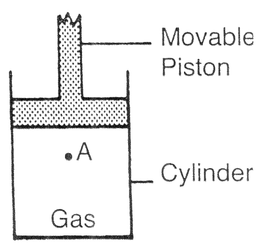
6. Which sample of methane gas contains the greatest number of molecules at standard temperature?

(1) 22.4 liters at 1 atmosphere
(2) 22.4 liters at 2 atmospheres
(3) 11.2 liters at 1 atmosphere
(4) 11.2 liters at 2 atmospheres

7. The pressure on 200. milliliters of a gas is decreased at constant temperature from 0.900 atm to 0.800 atm. The new volume of the gas, in milliliters, is equal to

(1) $200. \times \frac{0.900}{0.800}$
(2) $200. \times \frac{0.800}{0.900}$
(3) $0.800 \times \frac{200.}{0.900}$
(4) $0.800 \times \frac{0.900}{200.}$

8. The diagram below represents a gas confined in a cylinder fitted with a movable piston.



As the piston moves toward point *A* at a constant temperature, which relationship involving pressure (*P*) and volume (*V*) is correct?

(1) $P + V = k$ (2) $P - V = k$ (3) $P \div V = k$ (4) $P \times V = k$

9. What would be the volume at STP of 255 mL of CO_2 measured at 30.°C and 85.0 kPa?

(1) 193 mL (2) 1950 mL (3) 337 mL (4) 237 mL

10. Given: 400 mL of gas at STP. What is the pressure at 320 K if the volume is 500. mL?

(1) 0.683 atm (2) 0.938 atm (3) 1.07 atm (4) 1.47 atm

11. A gas occupies a volume of 560 mL at a temperature of 100°C. To what temperature must the gas be changed if it is to occupy 400 mL, with the pressure remaining unchanged?

(1) 71.4 K (2) 522 K (3) 100 K (4) 266 K

12. The pressure of a gas in a steel can is 92.0 kPa at 25°C. What will be the pressure if the temperature is increased to 50.°C?

(1) 46 kPa (2) 184 kPa (3) 99.7 kPa (4) 84.9 kPa

MIXTURES

Matter can exist as pure substances or mixtures. Mixtures are composed of two or more different substances that can be separated by physical means. The proportions of components in a mixture can be varied. Each component retains its original properties.

The physical separation of the components of the mixture depends on the differences in their physical properties such as density, practicle size, molecular polarity, boiling point, freezing point, and solubility. Some of the processes used to separate mixtures are described below.

Filtration. A mixture is poured through a porous material such as paper or fabric. It relies on the different sizes of the components in the mixture to make a separation. The smaller, usually molecular, component passes through the pores, but the larger components cannot. For example, if muddy water is poured into filter paper, the water molecules pass though, but the mud particles are caught in the paper.

Distillation. This method relies on the different boiling points of the mixture's components. The materials, which must be solids or liquids, are heated in a distilling flask. The lowest boiling point component boils first, leaving the other components in the flask. The vapor of the first component can be captured and condensed back into a liquid.