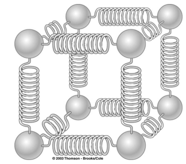
Thermal Physics

Essential Idea: Thermal physics deftly demonstrates the links between the macroscopic measurements essential to many scientific models with the microscopic properties that underlie these models.

What types of energy does the soccer ball have?



Macroscopic:

Microscopic:

Internal Kinetic Energy: arises from random translational, vibrational, and rotational motion of particles

Internal Potential Energy: due to intermolecular forces and stored in bonds

Internal Energy: **Symbol:** **Units:**

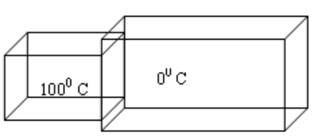
Temperature (Definition #1): **Symbol:** **Units:**

Celsius ↔ Kelvin conversion

1. Which has a greater . . .?

Quantity	Cup of Coffee	Iceberg
Temperature		
Average kinetic energy of molecules		
Total kinetic energy of molecules		
Internal energy		

2. What can you tell from the temperatures?



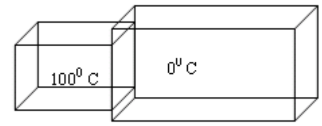
Thermal Energy: **Symbol:** **Units:**

Temperature (Definition #2):

Thermal Equilibrium:

Thermal Capacity and Specific Heat Capacity

What factors does the final equilibrium temperature depend on?



Thermal Capacity:

Formula: Symbol: Units:

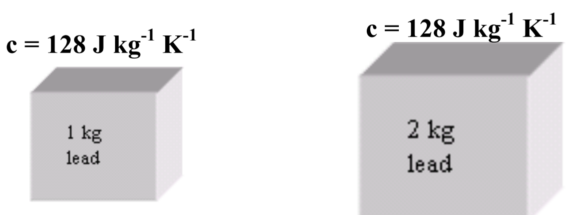
Specific Heat Capacity:

Formula: Symbol: Units:

Relationship:

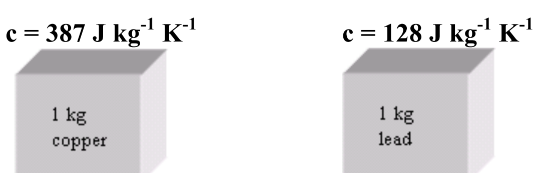
1. Compare the thermal capacities and specific heat capacities of these samples.

A



Why do different amounts of the same substances have different thermal capacities?

B



Why do the same amounts of different substances have different specific heat capacities?

2. One kilogram each of copper and lead are heated at the same constant rate on a hot plate.

   a) Which material will change its temperature the quickest?

   b) Which material will have the most internal energy if they are both heated to 90^0 C?

3. Substances with a high specific heat capacity . . .

1. The thermal capacity of a sample of lead is 3.2×10^3 J K^{-1}.

   a) How much thermal energy will be released if it cools from 61^0 C to 25^0 C?

   b) What is the mass of the sample?

Specific heat capacities

Substance	c_p (J/kg•°C)
aluminum	8.99×10^2
copper	3.87×10^2
glass	8.37×10^2
gold	1.29×10^2
ice	2.09×10^3
iron	4.48×10^2
lead	1.28×10^2
mercury	1.38×10^2
silver	2.34×10^2
steam	2.01×10^3
water	4.186×10^3

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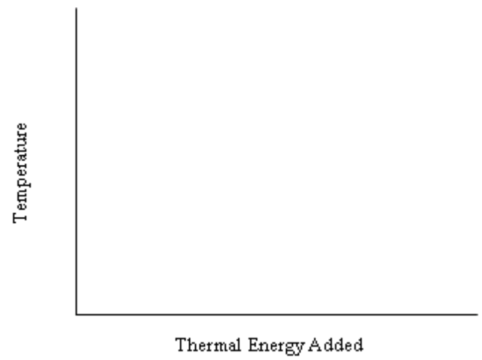
2. How much thermal energy is needed to raise the temperature of 2.50 g of water from its freezing point to its boiling point?



Slope

Compare your answer to the amount of thermal energy needed to raise the temperature of liquid mercury the same amount.

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3. A 50. watt heater is run for one minute to raise the temperature of a sample of lead from 310 K to 335 K. What is the mass of the sample?

4. An active solar heater is used to heat 50 kg of water initially at 12^0 C. If the average rate that the thermal energy is absorbed in a one hour period is 2.1×10^4 J min^{-1}, determine the final temperature after one hour.

Why will the final temperature probably be less than what you calculated above?

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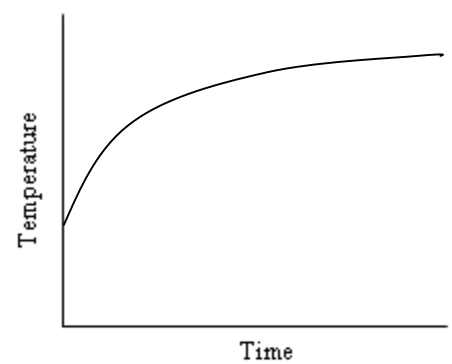
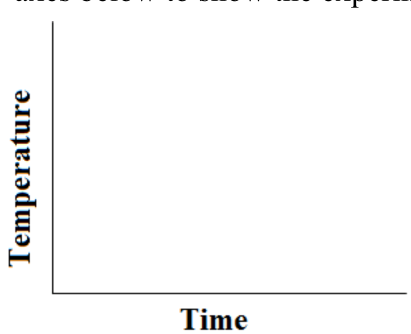
5. A hole is drilled in an 800g iron block and an electric heater is placed inside. The heater provides thermal energy at a constant rate of 600 W.

a) Assuming no thermal energy is lost to the surrounding environment, calculate how long it will take the iron block to increase its temperature by 15^0 C.

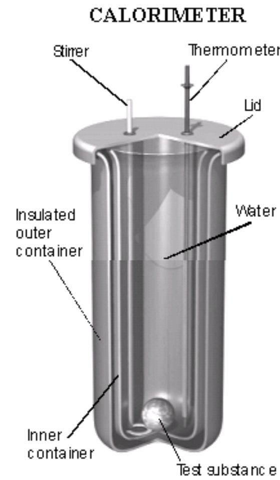
b) The temperature of the iron block is recorded as it varies with time and is shown at right. Comment on reasons for the shape of the graph.



c) The heater is removed and the water is allowed to cool. Sketch a line on the axes below to show the experimental relationship. Explain its shape.



Calorimetry



Method of Mixtures

Conservation of Energy

Assumption:

1. A 0.10 kg sample of an unknown metal is heated to 100^0 C by placing it in boiling water for a few minutes. Then it is quickly transferred to a calorimeter containing 0.40 kg of water at 10^0 C. After thermal equilibrium is reached, the temperature of the water is 15^0 C.

   a) What is the specific heat capacity of the metal sample?

   b) What is the thermal capacity of the metal sample?

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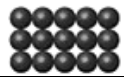
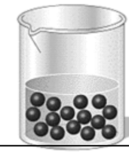
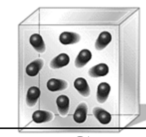
2. A 3.0 kg block of copper at 90^0 C is transferred to a calorimeter containing 2.00 kg of water at 20^0 C. The mass of the calorimeter cup, also made of copper, is 0.210 kg. Determine the final temperature of the water.

Phases of Matter

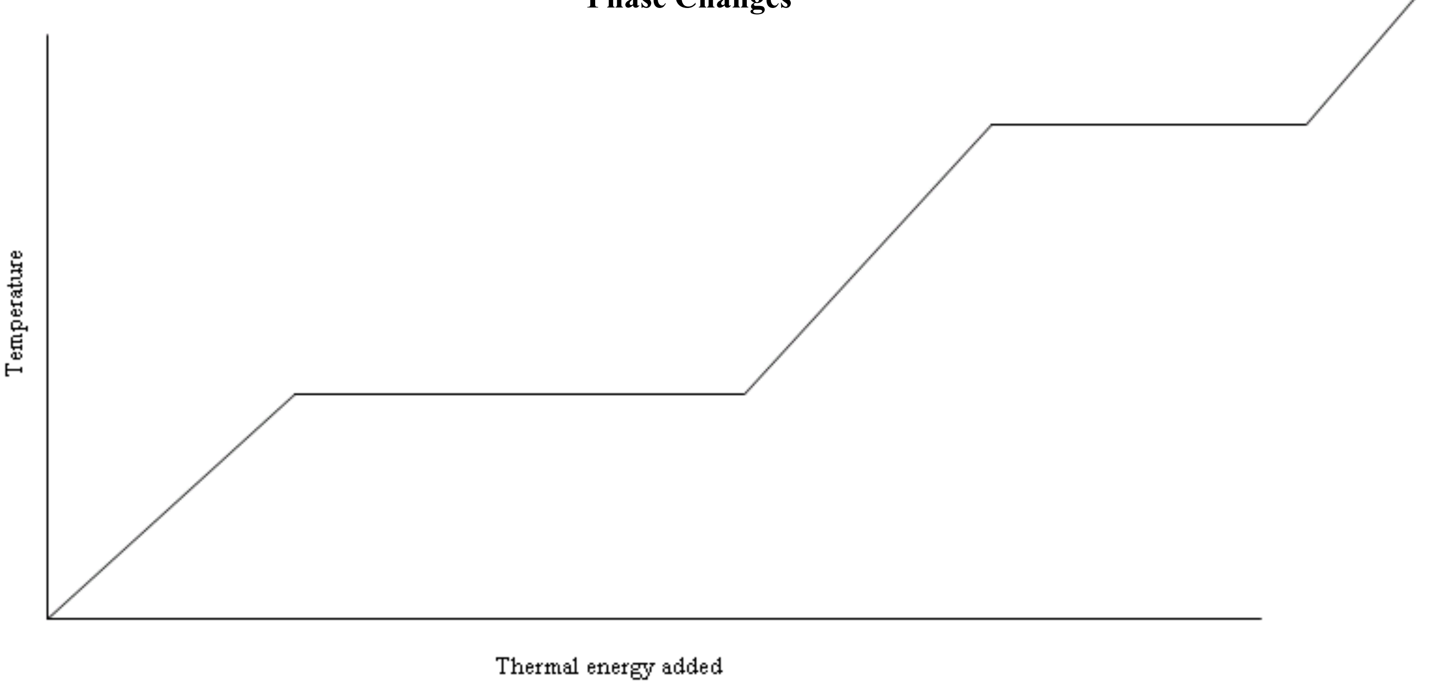
IB 12

Kinetic theory says that:
1. All matter is made up of atoms, and
2. the atoms are in continuous random motion at a variety of speeds.
3. Whether a substance is a solid, liquid, or gas basically depends on how close together its molecules are and how strong the bonds are that hold them together.

	Solid	Liquid	Gas
Macroscopic description	Definite volume Definite shape	Definite volume Variable shape	Variable volume Variable shape
Microscopic description	Molecules are held in fixed positions relative to each other by strong bonds and vibrate about a fixed point in the lattice	Molecules are closely packed with strong bonds but are not held as rigidly in place and can move relative to each other as bonds break and reform	Molecules are widely spaced apart without bonds, moving in random motion, and intermolecular forces are negligible except during collisions
Comparative density	High	High	Low
Density of . . .		Water =	Air =
Kinetic energy	Vibrational	Vibrational Rotational Some translational	Mostly translational Higher rotational Higher vibrational
Volume of molecules/ volume of substance	1	1	10^{-3}
Potential energy	High	Higher	Highest
Average molecular separation	Atomic radius	Atomic radius	10 x atomic radius
Molecules per m^3	10^{28}	10^{28}	10^{25}

Phase Changes



1. Describe and explain the process of phase changes in terms of molecular behavior.

When thermal energy is added to a solid, the molecules gain kinetic energy as they vibrate at an increased rate. This is seen macroscopically as an increase in temperature. At the melting point, a temperature is reached at which the kinetic energy of the molecules is so great that they begin to break the permanent bonds that hold them fixed in place and begin to move about relative to each other. As the solid continues to melt, more and more molecules gain sufficient energy to overcome the intermolecular forces and move about so that in time the entire solid becomes a liquid. As heating continues, the temperature of the liquid increases due to an increase in the vibrational, translational and rotational kinetic energy of the molecules. At the boiling point, a temperature is reached at which the molecules gain sufficient energy to overcome the intermolecular forces that hold them together and escape from the liquid as a gas. Continued heating provides enough energy for all the molecules to break their bonds and the liquid turns entirely into a gas. Further heating increases the translational kinetic energy of the gas and thus its temperature increases.

2. Explain in terms of molecular behavior why temperature does not change during a phase change.

The making or breaking of intermolecular bonds involves energy. When bonds are broken (melting and vaporizing), the potential energy of the molecules is increased and this requires input energy. When bonds are formed (freezing and condensing), the potential energy of the molecules is decreased as energy is released. The forming or breaking of bonds happens independently of the kinetic energy of the molecules. During a phase change, all energy added or removed from the substance is used to make or break bonds rather than used to increase or decrease the kinetic energy of the molecules. Thus, the temperature of the substance remains constant during a phase change.

3. Explain in terms of molecular behavior the process of evaporation.

Evaporation is a process by which molecules leave the surface of a liquid, resulting in the cooling of the liquid. Molecules with high enough kinetic energy break the intermolecular bonds that hold them in the liquid and leave the surface of the substance. The molecules that are left behind thus have a lower average kinetic energy and the substance therefore has a lower temperature.

Factors affecting the rate of evaporation:

a) surface area b) drafts c) temperature d) pressure e) latent heat of vaporization

4. Distinguish between evaporation and boiling.

Evaporation – process whereby liquid turns to gas, as explained above

- occurs at any temperature below the boiling temperature
- occurs only at surface of liquid as molecules escape
- causes cooling of liquid

Boiling – process whereby liquid turns to gas when the vapor pressure of the liquid equals the atmospheric pressure of its surroundings

- occurs at one fixed temperature, dependent on substance and pressure
- occurs throughout liquid as bubbles form, rise to surface and are released
- temperature of substance remains constant throughout process

Specific Latent Heat

Specific Latent Heat:

Symbol: Units:

Formula:

Specific latent heat of fusion:

Specific latent heat of vaporization:

Latent heats of fusion and vaporization at standard pressure

Substance	L_f (J/kg)	L_v (J/kg)
nitrogen	2.55×10^4	2.01×10^5
oxygen	1.38×10^4	2.13×10^5
ethyl alcohol	1.04×10^5	8.54×10^5
water	3.33×10^5	2.26×10^6
lead	2.45×10^4	8.70×10^5
aluminum	3.97×10^5	1.14×10^7

1. How much energy is needed to change 500 grams of ice into water?

   a) Assume the ice is already at its melting point.

   b) Assume the ice is at -15^0 C.

2. Thermal energy is supplied to a pan containing 0.30 kg of water at 20^0 C at a rate of 400 W for 10 minutes. Estimate the mass of water turned into steam as a result of this heating process.

3. The latent heat of vaporization of water is 2300 kJ/kg. How long would it take a 2 kW electric kettle containing 800g of boiling water to boil off all the water?

4. In order to maintain a constant body temperature, a sunbather need to lose about 320 J of thermal energy to the environment every second through sweating. Estimate the amount of sweat evaporated from the skin of the sunbather every hour.

5. The cost of electricity is \$0.15 per kWhr. How much does it cost to heat 1.0 m^3 of water from 20^o C to 25^oC?

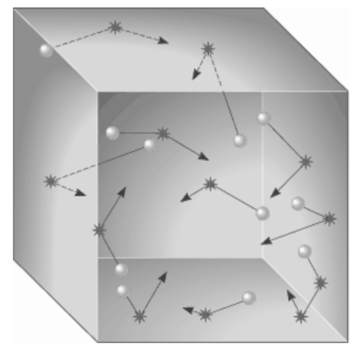
The Kinetic Model of an Ideal Gas

Essential Idea: The properties of ideal gases allow scientists to make predictions of the behaviour of real gases.

Kinetic theory views all matter as consisting of individual particles in continuous motion in an attempt to relate the macroscopic behaviors of the substance to the behavior of its microscopic particles.

Certain microscopic assumptions need to be made in order to deduce the behavior of an ideal gas, that is, to build the **Kinetic Model of an Ideal Gas**.

Assumptions of the Ideal Gas Model:



1. A gas consists of an extremely large number of very tiny particles (atoms or molecules) that are in continuous random motion with a variety of speeds.

2. The volume of the particles is negligible compared to the volume occupied by the entire gas.

3. The size of the particles is negligible compared to the distance between them.

4. Collisions between particles and collisions between particles and the walls of the container are assumed to be perfectly elastic and take a negligible amount of time. The time for a collision is much less than the time between collisions.

   Perfectly elastic:

5. In between collisions, the particles obey Newton's laws of motion, that is, they travel in straight lines at a constant speed.

6. No forces act between the particles except when they collide (no intermolecular forces).

Consequences of the Ideal Gas Model assumptions:

a)

b)

Compare real gases to an ideal gas:

a)

b)

The "state" of an ideal gas is described by the values of its pressure, volume, temperature, and amount.

Amount

N =

Avogadro's constant: the number of atoms in 12 g of carbon 12.

N_A =

Mole: an amount of a substance that contains as many particles as there are atoms in 12 grams of carbon-12.

Molar mass:

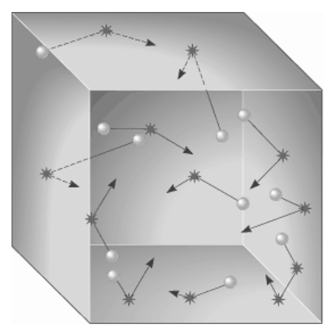
n = **Formula:**

1. As a general rule, the molar mass in grams of a substance is numerically equal to its mass number.

   a) 1 mole of $^{7}_{3}Li$ has a mass of

   b) 2 moles of $^{27}_{13}Al$ has a mass of

   c) How many atoms are in 8 grams of helium (He-4)?

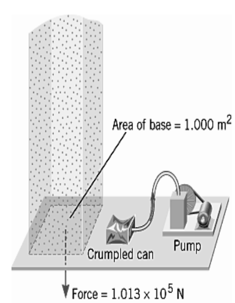


Pressure

Macroscopic definition: force per unit area acting on a surface

Microscopic definition: total force per unit area from the collisions of gas particles with walls of container

Formula: **Units:**

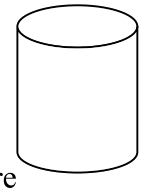


Atmospheric Pressure:

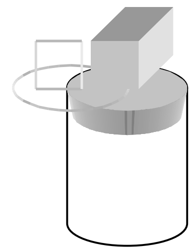
Atmospheric pressure at sea level

2. A cylinder with diameter 3.00 cm is open to the air.

   a) What is the pressure on the gas in this open cylinder?

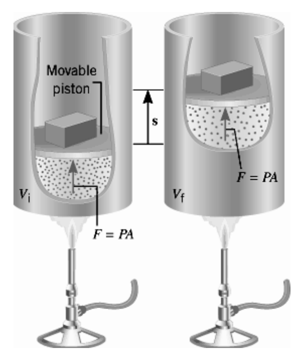


   b) What is the pressure on the gas after a 500. gram piston and a 5.00 kg block are placed on top?



Changes of State for an Ideal Gas

I. **Macroscopic behavior**: At a constant pressure, ideal gases increase in volume when their temperature increases.



Type of process:

Microscopic explanation: A higher temperature means faster moving particles that collide with the walls more often and with greater force. However, if the volume of the gas is allowed to increase, the rate at which these particles hit the walls will decrease and thus the average force exerted on the walls by the particles, that is, the pressure can remain the same.

Relationship:

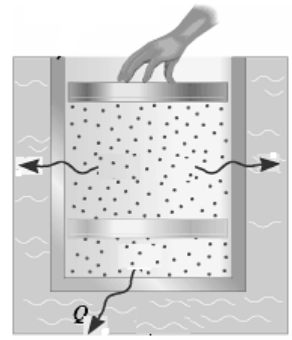
II. **Macroscopic behavior**: At a constant volume, ideal gases increase in pressure when their temperature increases.

Type of process:

Microscopic explanation: The increased temperature means the particles have, on average, more kinetic energy and are thus moving faster. This means that the particles hit the walls more often and, when they do, they exert a greater force on the walls during the collision. For both these reasons, the total force on the wall in a given time increases which means that the pressure increases.

Relationship:

III. **Macroscopic behavior**: At a constant temperature, ideal gases increase in pressure when their volume decreases.



Type of process:

Microscopic explanation: The decrease in volume means the particles hit a given area of the wall more often. The force from each particle remains the same but an increased number of collisions in a given time means the pressure increases.

Relationship:

Derivation:

Gas constant:

Equation of State:

Ideal Gas:

Derivation:

Combined Gas Law:

Assumption:

3. What is the volume occupied by 32 g of oxygen (mass number = 16) at room temperature and atmospheric pressure?

4. A gas in a closed container is under a pressure of 1 atm and a temperature of -173 ^{o}C. The gas is then heated to 27 ^{o}C. What is the new pressure of the gas?

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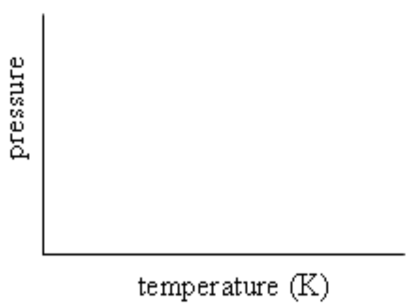
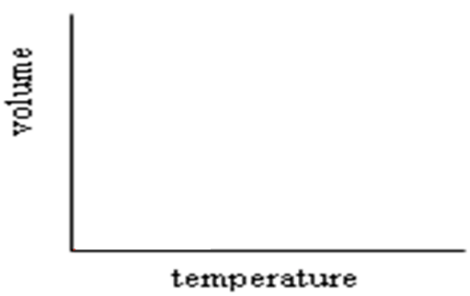
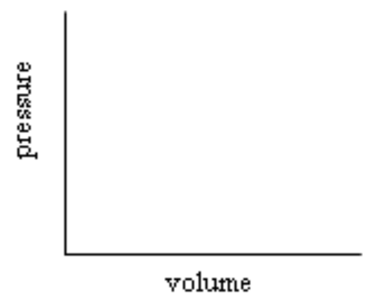
5. Sketch the graph of each relationship shown below for an ideal gas, state the control variable and the type of process.







Control =

Type of process:

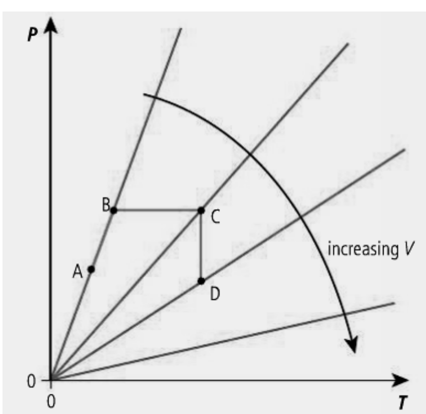
Control =

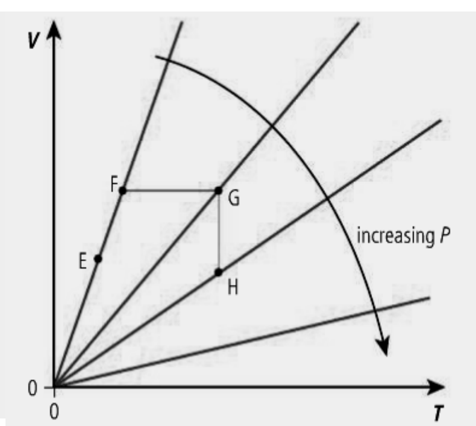
Type of process:

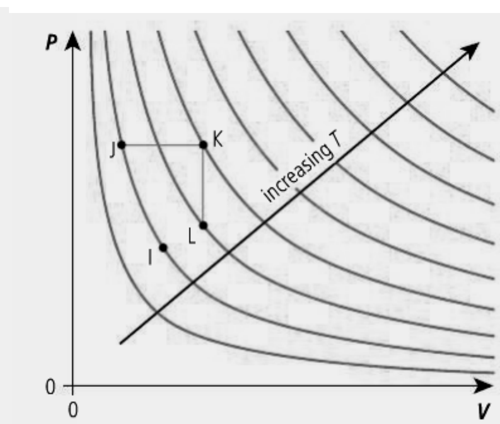
Control =

Type of process:

6. For each graph below, identify each of the indicated processes.







Process AB:

Process BC:

Process CD:

Process EF:

Process FG:

Process GH:

Process IJ:

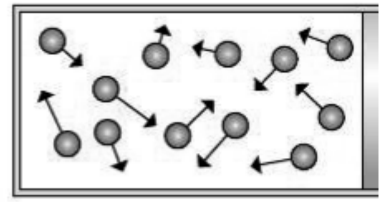
Process JK:

Process KL:

Internal Energy of a Monatomic Ideal Gas

Monatomic gas:

According to the assumptions of the kinetic model of an ideal gas, there are no bonds or interatomic forces between any of the particles in the gas.



Therefore,

a)

b)

c)

Formulas:

Derivation:

Where k_B=

1. A container holds 8.0 grams of helium gas (He-4) at 20° C.

   a) Calculate the average kinetic energy of the atoms.

   b) Calculate the total internal energy of the gas.

2. A sample of helium gas (He-4) and a sample of argon gas (Ar-40) are held at the same temperature in separate containers, as shown below. Compare the average speeds of the atoms.