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Physics of Heat

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Preface

Heat is one of the fundamental concepts in the field of thermodynamics, playing a pivotal role in our understanding of energy transfer and transformation. In this work, we delve into the nature of heat, exploring its properties, measurement, and the underlying principles that govern its behavior. With a focus on the connections between heat and the laws of thermodynamics, we aim to provide a comprehensive overview of this essential subject.

At the heart of our discussion lies the thermometer, a fascinating instrument that has revolutionized our ability to quantify heat. From the first rudimentary designs to the sophisticated digital devices of today, thermometers enable us to measure temperature—an intrinsic property of matter that directly relates to thermal energy. By understanding how thermometers work, we can appreciate the importance of accurate measurements in both scientific research and practical applications.

As we explore the concept of specific heat, we will uncover why different substances respond variably to the addition of heat. Specific heat capacity, a measure of how much heat is required to change the temperature of a unit mass of a substance, is crucial for understanding phenomena ranging from cooking to climate science. By studying specific heat, we gain insights into the thermal properties of materials and the energy exchanges that occur in natural and engineered systems.

Moreover, we will discuss the First Law of Thermodynamics, a cornerstone of physical science that encapsulates the principle of conservation of energy. This law asserts that energy cannot be created or destroyed; rather, it can only change forms. Through examples and real-world applications, we will illustrate how heat is transferred between systems and its role in conversions between thermal energy, mechanical work, and other forms of energy.

Throughout this book, we will emphasize the interconnectedness of these concepts, illustrating how they collectively contribute to a deeper understanding of the natural world. From the workings of engines to the intricacies of climate systems, the principles of heat and thermodynamics are intrinsic to many aspects of our lives. As we embark on this exploration, we invite readers to consider the profound implications of heat, not only as a scientific concept but also as a vital element that shapes our environment and experiences.

Join us as we uncover the mysteries of heat, enhance our appreciation for the laws governing energy, and empower ourselves with the knowledge that will spark further inquiry into the wonders of thermodynamics.

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Heat Measurements

Introduction

The study of heat transformations into other forms of energy, called thermodynamics, began with the eighteenth-century engineers who built the first steam engines. These steam engines were used to power trains, factories, and water pumps for coal mines, and thus they contributed greatly to the Industrial Revolution in Europe and in the United States. In learning to design more efficient engines, the engineers developed new concepts about how heat is related to useful work. Although the study of thermodynamics began in the eighteenth century, it was not until around 1900 that the concepts of thermodynamics were linked to the motions of atoms and molecules in solids, liquids, and gases.

Today, the concepts of thermodynamics are widely used in various applications that involve heat and temperature. Engineers use the laws of thermodynamics to continually develop higher performance refrigerators, automobile engines, aircraft engines, and numerous other machines.

Temperature

It is the measure of the degree of heat or coldness of an object based on a certain temperature scale. Heat always flows from a hot thing to a nearby cool object. Objects respond differently to heat based on their state and composition. For instance, a steel rod expands in length as its temperature rises. Thermometers are instruments used to measure temperatures.

Thermometric scale

When creating a thermometer, the following factors must be taken into account:

- 1- Choose a thermometer material whose physical properties change with temperature, such as the expansion of liquids or gases or the resistance of a platinum wire, etc.
- 2- It is necessary to choose two known constant temperatures, usually the melting point of ice (lower point) and the boiling point of water (upper point), at normal atmospheric pressure (76 cm Hg).
- 3- The difference between these two temperatures is divided into a number of equal sections, each of which is called a degree, and the interval in degrees between the two fixed points is called the fundamental interval. See figure (1).

Ch1: Temperature and Thermal Energy

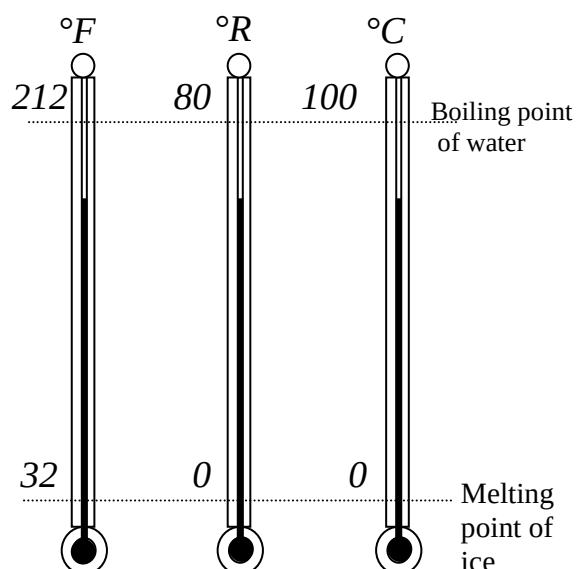


Figure (1)

The Celsius scale:

The scale is divided into 100 equal portions from 0°C to 100°C, named after the physicist Celsius.

Reaumur Scale:

The scale is broken into 80 divisions with the base period ranging from 0°R to 80°R.

Fahrenheit scale

This scale divides the base period into 180 equal divisions, with 32°F at the bottom point and 212°F at the higher point.

Relationship between thermometers:

A comparison of these systems can be found in the table below.

Fahrenheit scale °F	Rheometer scale °R	Celsius scale °C	Fixed points
32	zero	zero	Melting Point
212	80	100	Boiling Point

Table (1) Relationship between thermometric scales

With the help of the following formula, you will be able to convert from one of these scales

to the other: $\frac{C}{100} = \frac{F-32}{180} = \frac{R}{80} \rightarrow \frac{F-32}{C} = \frac{9}{5} \rightarrow F = \frac{9}{5}C + 32$

Example (1) determine the temperature of a container when the Fahrenheit reading is double the Celsius reading, and find the temperature at which the two readings are the same.

Solution

First, since $F=2C$ in this $F = \frac{9}{5}C + 32 \rightarrow 2C = \frac{9}{5}C + 32$. By substituting

$$2C - \frac{9}{5}C = 32 \rightarrow \frac{10-9}{5}C = 32 \rightarrow \frac{1}{5}C = 32 \rightarrow C = 160^{\circ}\text{C} \rightarrow F = 320^{\circ}\text{F}$$

$$\text{Also, } F = C \rightarrow C = \frac{9}{5}C + 32 \rightarrow C - \frac{9}{5}C = 32 \rightarrow \frac{5-9}{5}C = 32$$

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$$\therefore \frac{-4}{5} C = 32 \rightarrow C = -40^{\circ}\text{C} \rightarrow F = -40^{\circ}\text{F}$$

Example (2). If a Celsius thermometer reads 35 degrees Celsius, what will a Fahrenheit thermometer read in the same room? What is the Fahrenheit temperature on a day with a temperature of -10 degrees Celsius?

$$\text{At } C = 35^{\circ}\text{C we found } F = \frac{9}{5} C + 32 = \frac{9}{5} \times 35 + 32 = 95^{\circ}\text{F}$$

$$\text{In the second case } C = -10^{\circ}\text{C} \rightarrow F = \frac{9}{5} C + 32 = \frac{9}{5} (-10) + 32 = 14^{\circ}\text{F}$$

Absolute (Kelvin) scale

Absolute zero is defined as the temperature at which the volume of a gas theoretically vanishes with constant pressure. The degree of absolute zero on the Celsius scale is -273°C and can be converted from Celsius to absolute temperature (Kelvin) using the equation $K = C + 273$, meaning that the absolute scale zero is 273° below the Celsius scale zero itself, and the two scales are divided in the same way.

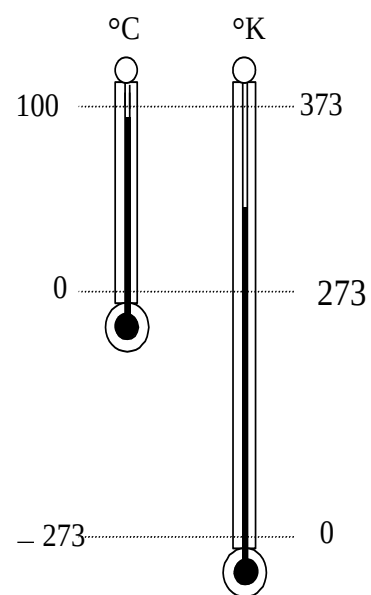


Figure (2) Absolute Scale and Celsius scale

Thermometers

Mercury thermometer

Mercury is one of the most suitable liquids for making thermometers, as it freezes at -40°C and boils at 357°C . It expands regularly and is opaque, so it can be easily seen through the glass and does not stick to the glass wall. It then takes the temperature of what it touches from the objects quickly, and its expansion coefficient is relatively large.



Figure (3) mercury thermometer

The thermometer consists of a thick-walled capillary tube of uniform cross-section, terminating at the bottom with a glass reservoir holding mercury. The tube is sealed at the top and evacuated from the air. Mercury is

commonly used in medical thermometers that are used to measure human temperature, so it

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is only listed from 35°C to 42°C or from 95f to 110f. It is characterized by a slight bend in the thermometer tube just above the mercury reservoir, so the mercury passes through this bend when the temperature rises, while it cannot return to the reservoir unless we shake the thermometer.

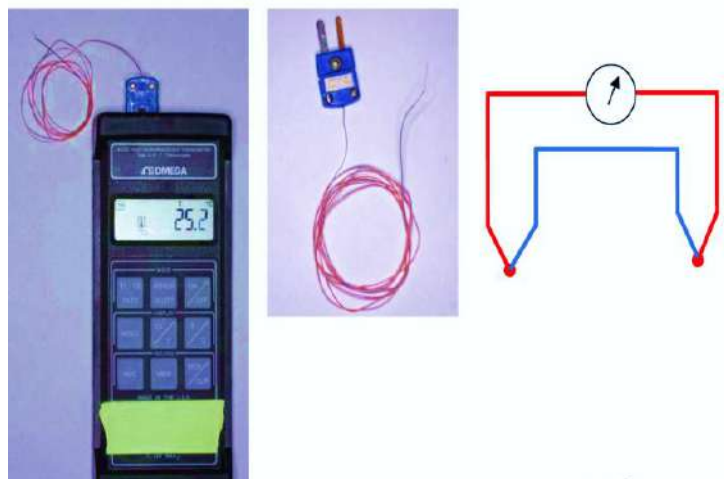
The alcohol thermometer

It can measure temperatures as low as -110°C. Alcohol is advantageous due to its greater expansion coefficient compared to mercury, making it more sensitive and reducing the impact of irregularities in the thermometer tube. This allows for a smaller reservoir in the thermometer tube. However, its non-uniform expansion makes it unsuitable for precise measurements, but it is commonly used in meteorology to determine atmospheric temperature.

Alcohol is utilized as a thermometer in place of mercury for measuring low temperatures due to its ability to remain in a liquid condition within the temperature range of -110°C to 78°C.

Thermocouple

When two different metals, like iron and copper, are connected to create a thermocouple, an electromotive force is produced as one joint's temperature increases compared to the other. The potential difference is determined by the temperature difference between the two joints. When the cold junction is immersed in crushed ice to maintain a temperature of zero degrees Celsius and the other junction is placed in a hot medium, the galvanometer deflects in direct



proportion to the Celsius temperature of the hot junction. Calibrate the galvanometer by placing the hot junction in water boiling at 100 degrees Celsius and dividing the deflection magnitude into 100 divisions, with each division representing a degree Celsius.

Expansion of solids

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When the temperature of an object increases, the kinetic energy of its molecules rises, causing an increase in the amplitude of their vibration. This results in an expansion of the object due to the increased average distance between its molecules. The expansion will occur uniformly in all directions for the body. However, if the body is shaped like a long rod, the expansion in length will be more prominent, and the expansion in smaller dimensions can be disregarded. In this scenario, we focus on longitudinal expansion. If the body is in the form of a thin slice, expansion in a thin slice results in superficial expansion, where expansion in the thickness of the slice is negligible. In contrast, a body in the form of a mold experiences volumetric expansion, expanding in all directions.

Longitudinal expansion of solids:

It was found experimentally that if a metal wire of original length L_1 is heated from temperature T_1 to temperature T_2 , it will expand, and the resulting elongation of the wire is directly proportional to its temperature rise ($T_2 - T_1$) as well as directly proportional to the original length of the wire, i.e.

$$\Delta T = T_2 - T_1 \rightarrow \Delta L \propto L_1 \rightarrow \Delta L = L_2 - L_1 = \alpha L_1 (T_2 - T_1)$$

$$L_2 = L_1 + \alpha L_1 (T_2 - T_1) \rightarrow L_2 = L_1 \{1 + \alpha (T_2 - T_1)\}$$

Where ΔL is the elongation of the wire, that is, its new length after stretching minus its original length before stretching, and α is the coefficient of longitudinal expansion of the wire in units of $^{\circ}\text{C}^{-1}$.

Special case: given that, the length of the wire at 0°C is L_0 and its length at T is L_T , then

$$L_T = L_0 \{1 + \alpha T\}$$

Table 2 shows the longitudinal coefficients of thermal expansion for some materials.

Example (3) a grinded and polished Pyrex glass column was 10 cm long when the room temperature was 20 °C. If the temperature of this column is raised to 420 °C, calculate

a) The amount of elongation. B) The new length of the wire after raising its temperature from 20 °C to 420 °C? Knowing that $\alpha = 3 \times 10^{-6} \text{ } ^\circ\text{C}^{-1}$

Solution

$$\Delta L = \alpha L_1 (T_2 - T_1) = 3 \times 10^{-6} \times 10 \times (420 - 20) = 0.012 \text{ cm}$$

$$L_2 = L_1\{1 + \alpha (T_2 - T_1)\} = 10\{1 + 3 \times 10^{-6} (420 - 20)\} = 10.012cm$$

Also, $L_2 = L_1 + \Delta L = 10 + 0.012 = 10.012cm$

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Example (4) a rod made of silver is exactly 50cm long at a temperature of 200°C Calculate its length when the temperature is 400°C, given that the coefficient of longitudinal expansion of silver is $\alpha=18 \times 10^{-6} \text{ }^{\circ}\text{C}^{-1}$, and then compute the elongation caused by this temperature change.

Solutioin:

$$\Delta L = \alpha L_1 [T_2 - T_1] = 18 \times 10^{-6} \times 50 \times (400 - 200)$$
$$\Delta L = 0.18 \text{ cm} \rightarrow L_2 = L_1 + \Delta L = 50 + 0.18 = 50.18 \text{ cm}$$

Table 2

Average Expansion Coefficients for Some Materials Near Room Temperature			
Material	Average Linear Expansion Coefficient (α)($^{\circ}\text{C}$) $^{-1}$	Material	Average Volume Expansion Coefficient (β)($^{\circ}\text{C}$) $^{-1}$
Aluminum	24×10^{-6}	Alcohol, ethyl	1.12×10^{-4}
Brass and bronze	19×10^{-6}	Benzene	1.24×10^{-4}
Copper	17×10^{-6}	Acetone	1.5×10^{-4}
Glass (ordinary)	9×10^{-6}	Glycerin	4.85×10^{-4}
Glass (Pyrex)	3.2×10^{-6}	Mercury	1.82×10^{-4}
Lead	29×10^{-6}	Turpentine	9.0×10^{-4}
Steel	11×10^{-6}	Gasoline	9.6×10^{-4}
Invar (Ni-Fe alloy)	0.9×10^{-6}	Air ^a at 0°C	3.67×10^{-3}
Concrete	12×10^{-6}	Helium ^a	3.665×10^{-3}

Thermal expansion of the Mackinac Bridge:

Example (5) The main span of the Mackinac Bridge (see the Figure) has a length of 1158 m. The bridge is built of steel. Assume that the lowest possible temperature experienced by the bridge is $-50 \text{ }^{\circ}\text{C}$ and the highest possible temperature is $50 \text{ }^{\circ}\text{C}$. How much room must be made available for thermal expansion of the center span of the Mackinac Bridge?

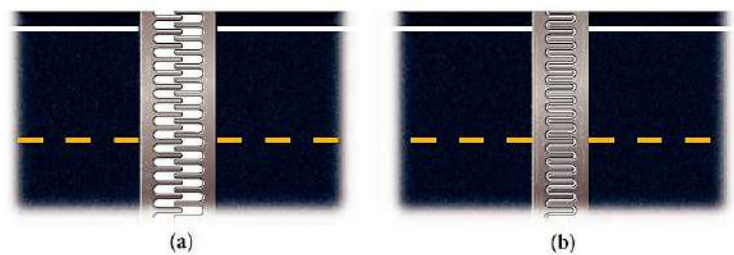


Solution: The linear expansion coefficient of steel is $\alpha = 13 \times 10^{-6} \text{ }^{\circ}\text{C}^{-1}$. Thus, the total linear expansion of the center span of the bridge that must be allowed for is given by

$$\Delta L = \alpha L \Delta T = 13 \times 10^{-6} \times 1158 \times (50 - (-50)) = 1.5 \text{ m}$$

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Discussion: A change in length of 1.5 m is fairly large. How is this length change accommodated in practice? (Obviously, we cannot have gaps in the road surface.) The answer lies in expansion joints, which are metal connectors between



Finger joints between road segments: (a) open and (b) closed.

bridge segments whose parts can move relative to each other. A popular type of expansion joint is the finger joint (see the Figure). The Mackinac Bridge has two large finger joints at the towers to accommodate the expansion of the suspended parts of the roadway and eleven smaller finger joints and five sliding joints across the main span of the bridge.

Area expansion

The effect of a change in temperature on the area of an object is analogous to using a copy machine to enlarge or reduce a picture. Each dimension of the object will change linearly with the change in temperature.

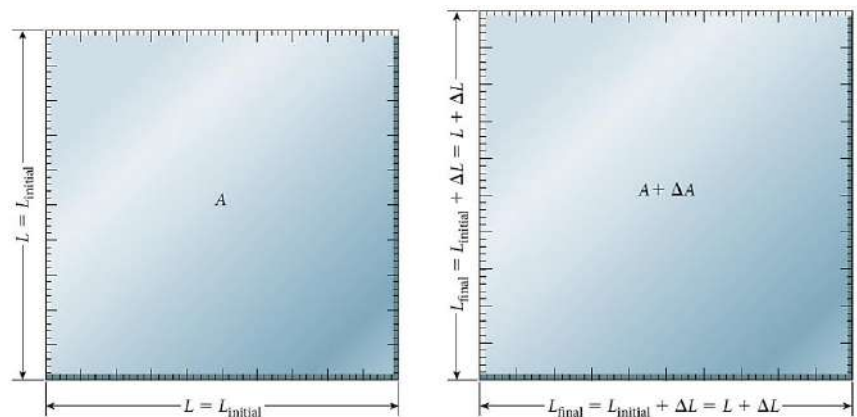


FIGURE Thermal expansion of a square plate with side L .

Quantitatively, for a square object with side L (see the Figure), the area is given by $A = L^2$. Taking the differential of both sides of this equation, we get $dA = 2L dL$. If we make the approximations that $\Delta A = dA$ and $\Delta L = dL$, we can write $\Delta A = 2L\Delta L$. Then we then obtain $\Delta A = 2L(\alpha L\Delta T) = 2\alpha A\Delta T$. Although a square was used to derive the last equation, it holds for a change in area of any shape.

Volume expansion

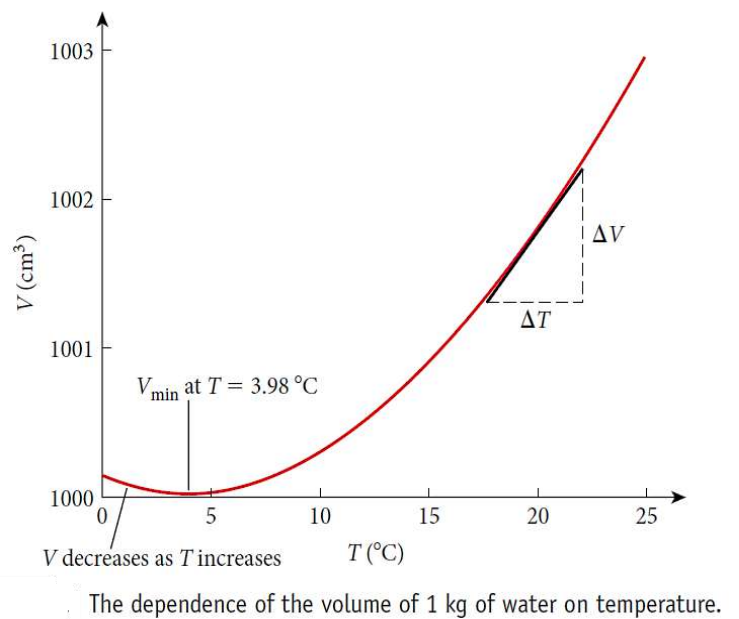
Now let us consider the change in volume of an object with a change in temperature. For a cube with side L , the volume is given by $V = L^3$. Taking the differential of both sides of this equation, we get $dV = 3L^2dL$. Making the

Table	Volume
Expansion Coefficient for Some Common Liquids	
Material	$\beta(10^{-6} \text{ }^\circ\text{C}^{-1})$
Mercury	181
Gasoline	950
Kerosene	990
Ethyl alcohol	1120
Water (1 °C)	-47.8
Water (4 °C)	0
Water (7 °C)	45.3
Water (10 °C)	87.5
Water (15 °C)	151
Water (20 °C)	207

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approximations that $\Delta V = dV$ and $\Delta L = dL$, we can write $\Delta V = 3L^2\Delta L$. Then, we obtain $\Delta V = 3L^2(\alpha L\Delta T) = 3\alpha V\Delta T$. Because the change in volume with a change in temperature is often of interest, it is convenient to define the volume expansion coefficient: $\beta = 3\alpha$. Thus, we can rewrite equation as, $\Delta V = \beta V\Delta T$. Although a cube was used to derive the equation, it can be generally applied to a change in the volume of any shape. Some typical volume expansion coefficients are listed in the pervious Table. This equation $\Delta V = \beta V\Delta T$ applies to the thermal expansion of most solids and liquids. However, it does not describe the thermal expansion of water. Between 0 °C and about 4 °C, water contracts as the temperature increases (see the oppsite Figure).

Water with a temperature above 4 °C is denser than water with a temperature just below 4 °C. This property of water has a dramatic effect on the way a lake freezes in the winter. As the air temperature drops from warm summer to cold winter



temperatures, the water in a lake cools from the surface down. The cooler, denser water sinks to the bottom of the lake. However, as the temperature of the water on the surface falls below 4 °C, the downward motion ceases, and the cooler water remains at the surface of the lake, with the denser, warmer water below. The top layer eventually cools to 0 °C and then freezes. Ice is less dense than water, so the ice floats on the water. This newly formed layer of ice acts as insulation, which slows the freezing of the rest of the water in the lake. If water had the same thermal expansion properties as other common materials, instead of freezing from the top down, the lake would freeze from the bottom up, with warmer water remaining at the surface of the lake and cooler water sinking to the bottom.

This would mean that lakes would freeze solid more often, and any forms of life in them that could not exist in ice would not survive the winter. In addition, you can see from Figure that the volume of a given amount of water never depends linearly on the temperature. However, the linear dependence of the volume of water on temperature can be

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approximated by considering a small temperature interval. The slope of the volume/temperature curve is $\Delta V/\Delta T$, so we can extract an effective volume expansion coefficient for small temperature changes.

For example, the volume expansion coefficient for water at six different temperatures is given in the last Table; note that at 1 °C, $\beta = -47.8 \times 10^{-6} \text{ } ^\circ\text{C}^{-1}$, which means that the volume of water will decrease as the temperature is increased.

Thermal expansion of gasoline

Example 6: You pull your car into a gas station on a hot summer day, when the air temperature is 40 °C. You fill your empty 55-L gas tank with gasoline that comes from an underground storage tank where the temperature is 12 °C. After paying for the gas, you decide to walk to the restaurant next door and have lunch. Two hours later, you come back to your car and discover that gasoline has spilled out of the gas tank onto the ground.

Problem: How much gasoline has spilled?

Solution: We know the following: The temperature of the gasoline you put in the tank starts out at 12 °C. The gasoline warms up to the outside air temperature of 40 °C. The volume expansion coefficient of gasoline is $950 \times 10^{-6} \text{ } ^\circ\text{C}^{-1}$.

While you were gone, the temperature of the gasoline changed from 12 °C to 40 °C. We can find the change in volume of the gasoline as the temperature increases using the equation: $\Delta V = \beta V \Delta T = (950 \times 10^{-6} \text{ } ^\circ\text{C}^{-1}) (55 \text{ L}) (40 \text{ } ^\circ\text{C} - 12 \text{ } ^\circ\text{C}) = 1.5 \text{ L}$.

Thus, the volume of the gasoline increases by 1.5 liter as the temperature of the gasoline is raised from 12 °C to 40 °C. The gas tank was full when the temperature of the gasoline was 12 °C, so this excess volume spills out of the tank onto the ground.

Is heat a form of energy? If you rub your hands against each other, you will notice that they warm up. As a result of moving your hands against each other against frictional forces, the kinetic energy you gave them is converted into heat. Since energy never goes away, we can

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say that the disappearance of kinetic energy and the appearance of heat can be explained on the sole basis that heat is a form of energy.

What happens if this heat energy is given to a substance?

Obviously, the substance will become hotter; that is, its temperature will rise, and it may melt or vaporize; that is, there may be a change of state.

Heat capacity of an object:

Assume we have a body and want to increase its overall temperature by one degree Celsius. The amount of heat given to the body is known as the body's heat capacity and is denoted by the symbol C . The unit of measurement for heat capacity in the c.g.s system is the calorie per degree, where $1\text{Cal}=4.186\text{ Joule}$.

Definition of the heat capacity of an object:

It is the amount of heat needed to raise the body temperature by one degree Celsius.

Where Q is the amount of heat, $C = \frac{\Delta Q}{\Delta T}$.

Specific heat of a substance:

The heat capacity of an object varies with its mass. In other words, this amount of heat is not characteristic of a substance and is not an attribute of it. So scientists thought about

choosing the amount of heat needed to raise the temperature of a unit mass of matter by one degree, and it called "specific heat capacity" or "specific heat" and is symbolized by the symbol S .

Definition of the specific heat of a substance

It is the amount of heat needed to raise the temperature of a unit mass of a substance by one degree, the formula for it is as follow, $= \frac{\Delta Q}{m \Delta T}$.

From the above equation, we can see that the unit for specific heat is $\text{Cal/gm}^\circ\text{C}$, for example, the specific heat of water and ice are respectively (Table 3 show the values of specific heat of

Table 3

Specific Heats of Some Substances at 25°C and Atmospheric Pressure

Substance	Specific heat c	
	$\text{J/kg} \cdot ^\circ\text{C}$	$\text{cal/g} \cdot ^\circ\text{C}$
<i>Elemental solids</i>		
Aluminum	900	0.215
Beryllium	1 830	0.436
Cadmium	230	0.055
Copper	387	0.092 4
Germanium	322	0.077
Gold	129	0.030 8
Iron	448	0.107
Lead	128	0.030 5
Silicon	703	0.168
Silver	234	0.056
<i>Other solids</i>		
Brass	380	0.092
Glass	837	0.200
Ice (-5°C)	2 090	0.50
Marble	860	0.21
Wood	1 700	0.41
<i>Liquids</i>		
Alcohol (ethyl)	2 400	0.58
Mercury	140	0.033
Water (15°C)	4 186	1.00
<i>Gas</i>		
Steam (100°C)	2 010	0.48

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some values at 15 °C) : $s_w = \frac{1 \text{ cal}}{\text{gm } ^\circ\text{C}} = \frac{4.186 \text{ J}}{\text{gm } ^\circ\text{C}}$, $s_{ice} = \frac{0.5 \text{ cal}}{\text{gm } ^\circ\text{C}} = \frac{2.09 \text{ J}}{\text{gm } ^\circ\text{C}}$.

If an object is given a quantity of heat Q to raise the temperature of m masses of matter by dT , then $Q = m S (T_f - T_i) = m S \Delta T$. Where Q is the amount of heat gained or lost in units of calories in the c.g.s system or joules in the m.k.s system. The above two relations can calculate the amount of heat gained or lost by a mass m of a substance when its temperature changes from T_i to T_f provided that the state of the substance does not change from one form to another. If the state of the substance changes from one form to another, the previous relationship can not be used during the transformation process because the temperature during this process will remain constant, and the amount of heat gained or lost during the transformation process can be calculated from the following relationship: $Q = mL$, Where L is the latent heat and its unit in the c.g.s system is Cal/gm.

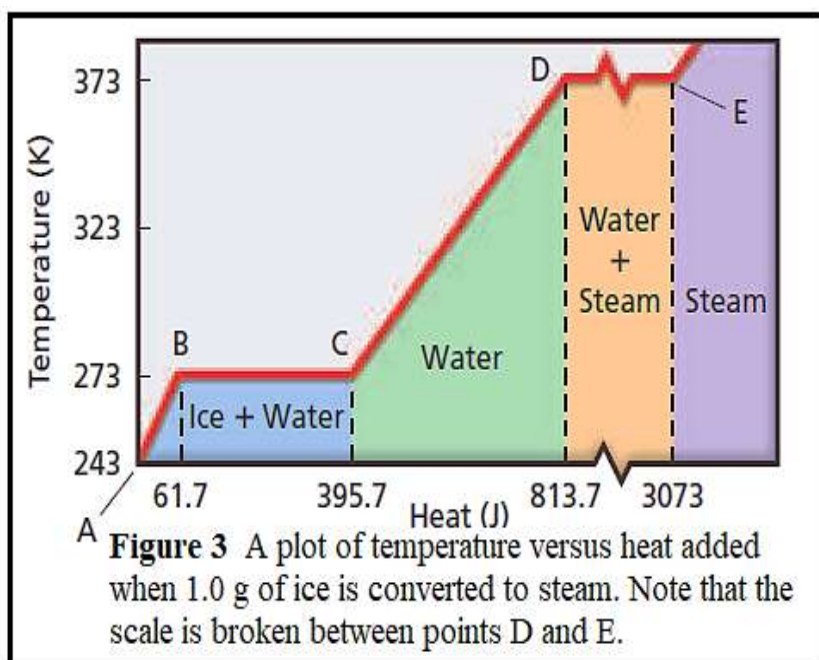
Changes of State and the Laws of Thermodynamics

Eighteenth-century steam engine builders used heat to turn liquid water into steam. The engine turned when the steam pushed a piston, cooled, and condensed back into a liquid. Adding heat to the liquid water changed not only its temperature but also its structure. You will learn that changing state means changing form as well as changing the way in which atoms store thermal energy.

Changes of State

The three most common states of matter are solids, liquids, and gases. As the temperature of a solid is raised, it usually changes to a liquid. At even higher temperatures, it becomes a gas. How can these changes be explained? Consider a material in a solid state. When the thermal energy of the solid is increased, the motion of the particles also increases, as does the temperature.

Figure 3 diagrams the changes of state as thermal energy is added to 1.0 g of water starting at 243 K (ice) and continuing until it reaches



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473 K (steam). Between points A and B, the ice is warmed to 273 K. At some point, the added thermal energy causes the particles to move rapidly enough that their motion overcomes the forces holding the particles together in a fixed location. The particles are still touching each other, but they have more freedom of movement. Eventually, the particles become free enough to slide past each other.

Melting point

At this point, the substance has changed from a solid to a liquid. The temperature at which this change occurs is the melting point of the substance. When a substance is melting, all of the added thermal energy goes to overcome the forces holding the particles together in the solid state. None of the added thermal energy increases the kinetic energy of the particles. This can be observed between points B and C in Figure 3, where the added thermal energy melts the ice at a constant 273 K. Because the kinetic energy of the particles does not increase, the temperature does not increase between points B and C.

Boiling point

Once a solid is completely melted, there are no more forces holding the particles in the solid state. Adding more thermal energy again increases the motion of the particles, and the temperature of the liquid rises. In the diagram, this process occurs between points C and D. As the temperature increases further, some particles in the liquid acquire enough energy to break free from the other particles. At a specific temperature, known as the boiling point, further addition of energy causes the substance to undergo another change of state. All the added thermal energy converts the substance from the liquid state to the gaseous state.

As in melting, the temperature does not rise while a liquid boils. In Figure 12-10, this transition is represented between points D and E. After the material is entirely converted to gas, any added thermal energy again increases the motion of the particles, and the temperature rises. Above point E, steam is heated to temperatures greater than 373 K.

Heat of fusion

The amount of energy needed to melt 1 kg of a substance is called the **heat of fusion** of that substance. For example, the heat of fusion of ice is 3.34×10^5 J/kg (80 Cal/gm). If 1 kg of ice at its melting point, 273 K, absorbs 3.34×10^5 J, the ice becomes 1 kg of water at the same temperature, 273 K. The added energy causes a change in state but not in temperature. The horizontal distance in Figure 3 from point B to point C represents the heat of fusion.

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Heat of vaporization

At normal atmospheric pressure, water boils at 373 K. The thermal energy needed to vaporize 1 kg of a liquid is called the **heat of vaporization**. For water, the heat of vaporization is $2.26 \times 10^6 \text{ J/kg}$ (540 Cal/gm). The distance from point D to point E in Figure 3 represents the heat of vaporization. Every material has a characteristic heat of vaporization. Between points A and B, there is a definite slope to the line as the temperature is raised. This slope represents the specific heat of the ice.

The slope between points C and D represents the specific heat of water, and the slope above point E represents the specific heat of steam. Note that the slope for water is less than those of both ice and steam. This is because water has a greater specific heat than does ice or steam. The heat, Q , required to melt a solid of mass m is given by the following equation

Heat Required to Melt a Solid $Q = mL_f$

The heat required to melt a solid is equal to the mass of the solid times the heat of fusion of the solid. Similarly, the heat, Q , required to vaporize a mass, m , of liquid is given by the following equation.

Heat Required to Vaporize a Liquid $Q = mL_v$

The heat required to vaporize a liquid is equal to the mass of the liquid times the heat of vaporization of the liquid.

When a liquid freeze, an amount of heat, $Q = -mL_f$, must be removed from the liquid to turn it into a solid. The negative sign indicates that the heat is transferred from the sample to the external world. In the same way, when a vapor condenses to a liquid, an amount of heat, $Q = -mL_v$, must be removed from the vapor.

Note that: When heat transferred from one object to another,

The amount of heat lost = the amount of heat gained

Example (7) what is the amount of heat required to convert a piece of ice of mass 720gm at a temperature of -10°C to a liquid state at a temperature of 15°C .

Knowing that $S_{ice} = 2220 \text{ J/kg.k}^0$, $L = 333 \times 10^3 \text{ J/Kg}$, $S_{water} = 4190 \text{ J/kg.k}^0$?

The solution

First stage: from ice at -10°C to ice at zero $^\circ\text{C}$

$$Q_1 = ms\Delta T = ms(T_f - T_i), \quad T_i = -10^\circ\text{C}, T_f = 0^\circ\text{C} \rightarrow$$

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$$Q_1 = 0.72 \times 2220(0 - (-10)) = 15984 J$$

Phase 2: Change from ice at zero °C to water at zero °C

$$Q_2 = mL = 0.72 \times 333 \times 10^3 = 239760 J$$

The third stage: from water at zero °C to water at 15 °C

$$Q_3 = ms\Delta T = ms(T_f - T_i), T_i = 0^\circ C, T_f = 15^\circ C \rightarrow$$

$$Q_3 = 0.72 \times 4190 (15 - 0) = 45252 J$$

$$Q = Q_1 + Q_2 + Q_3 = 300996 J$$

Table 4

Latent Heats of Fusion and Vaporization

Substance	Melting Point (°C)	Latent Heat of Fusion (J/kg)	Boiling Point (°C)	Latent Heat of Vaporization (J/kg)
Helium	-269.65	5.23×10^3	-268.93	2.09×10^4
Nitrogen	-209.97	2.55×10^4	-195.81	2.01×10^5
Oxygen	-218.79	1.38×10^4	-182.97	2.13×10^5
Ethyl alcohol	-114	1.04×10^5	78	8.54×10^5
Water	0.00	3.33×10^5	100.00	2.26×10^6
Sulfur	119	3.81×10^4	444.60	3.26×10^5
Lead	327.3	2.45×10^4	1750	8.70×10^5
Aluminum	660	3.97×10^5	2450	1.14×10^7
Silver	960.80	8.82×10^4	2193	2.33×10^6
Gold	1063.00	6.44×10^4	2660	1.58×10^6
Copper	1083	1.34×10^5	1187	5.06×10^6

Example (8)

A copper ball of mass $m=75\text{gm}$ is heated to a temperature $T=312^\circ\text{C}$. The ball is dropped into a beaker containing water of mass $m=220\text{ gm}$, the heat capacity of the beaker is $C=45\text{ cal/k}$ and the initial temperature of the water and glass is $T=12^\circ\text{C}$. Find the final temperature T_f of the copper ball, the beaker, and the water at thermal equilibrium, knowing that:

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$$s = 0.092 \text{ cal} / \text{g} \cdot \text{K}, s_w = 1 \text{ cal} / \text{g} \cdot \text{K}$$

$$m_{\text{copper}} = 75 \text{ gm}, T_{\text{copper}} = 312^\circ \text{C}$$

$$m_{\text{water}} = 220 \text{ gm}, C_{\text{beaker}} = 45 \text{ cal} / \text{K} T_i = 12^\circ \text{C}$$

The solution

The amount of heat gained = the amount of heat lost

$$Q_{\text{copper}} = Q_{\text{water}} + Q_{\text{beaker}}$$

$$m_c s_c (T_i - T_f) = m_w s_w (T_f - T_i) + c_b (T_f - T_i)$$

$$75 \times 0.092 (312 - T_f) = 220 \times 1 (T_f - 12) + 45 (T_f - 12)$$

$$2152.8 - 6.9 T_f = 220 T_f - 2640 + 45 T_f - 540$$

$$5332.8 = 271.9 T_f \rightarrow T_f = 19.61^\circ \text{C}$$

Transmission of heat

There are three different ways in which heat is transferred from one place to another: conduction, convection, and radiation.

Conduction

If you hold a metal rod at one end and then place the other end in a flame, you will soon feel the metal rod getting hot, indicating that heat has been transferred through it. On the other hand, if you hold a slice of wood by one end and then place the other end in the fire, the heat will not be transferred even a little bit, even if the end in the fire starts to ignite. In other words, metal is a good conductor of heat, while wood is a poor conductor of heat. The word "conduction" is used to mean heat transfer in solids.

Thermal conductivity is defined as follows: It is the process by which heat is transferred from one point to another through solids without moving the molecules of the material itself.

In the previous example, we find that when the end of the metal rod that is exposed to the flame is heated, the molecules of this end will vibrate capacitively from one molecule to another neighboring molecule, and so the process

continues until the heat is transferred from the hot end to the cold end.

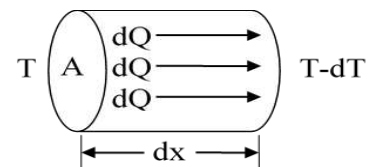
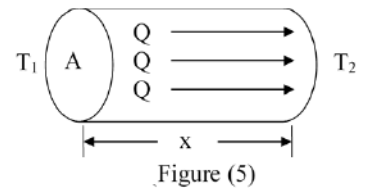


Figure (4)

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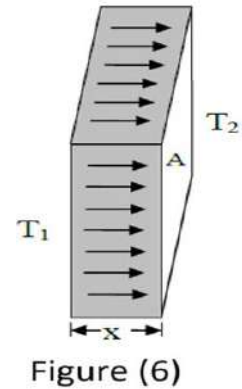
Suppose we have a section of a metal rod with thickness dx , cross-sectional area A , temperature at one end T and the other end $T-dT$ as shown in figure (4). The heat flow rate (the amount of heat flowing in a unit of time) is directly proportional to the cross-sectional area of the rod A and the difference in the temperatures between the two sides of the rod



$$\text{i.e. } \frac{dQ}{dt} \propto A (-dT),$$

And inversely proportional to the thickness dx i.e

$$\begin{aligned} \frac{dQ}{dt} &\propto \frac{1}{dx} \rightarrow \frac{dQ}{dt} \propto \frac{A (-dT)}{dx} \\ \therefore \frac{dQ}{dt} &= K \frac{A (-dT)}{dx} = -K \frac{A (T_2 - T_1)}{dx} \\ \therefore \frac{dQ}{dt} &= K \frac{A (T_1 - T_2)}{dx} \end{aligned}$$



Where K is the coefficient of thermal conductivity and its unit is $\text{cal/cm sec } ^\circ\text{C}$

The quantity dT/dx represents the rate of change of temperature with respect to distance, when the temperature decreases as the distance from the hot end of the rod increases, the quantity dT/dx is negative and is called the temperature gradient. See Figure (5).

In the steady state, when a metal rod of length x has temperatures T_1 and T_2 at its two ends, the difference between T_2 and T_1 remains constant. To prevent thermal energy loss, it is necessary to isolate the rod from the external medium. The amount of heat passing through the rod during time t can be calculated as follows: $Q = K \frac{A (T_1 - T_2)}{x} t$

Example (9) a metal slide is 0.2 cm thick and has a surface area of 200 cm^2 . If the temperature difference between the two opposite surfaces is 100°C , find the amount of heat that will transferred through the slide in a time of one minute, given that $k = 0.2 \text{ cal/sec.cm.}^\circ\text{C}$

The solution:

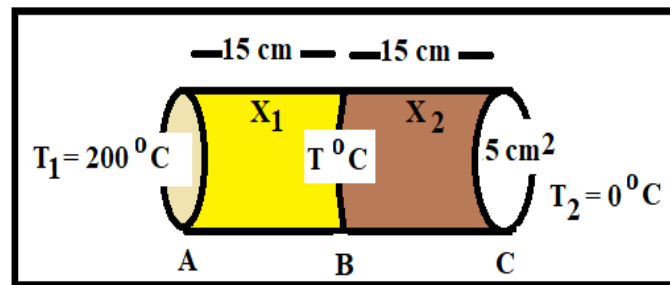
$$k = 0.2 \frac{\text{cal}}{\text{sec.cm.}^\circ\text{C}}, A = 200 \text{ cm}^2, x = 0.2 \text{ cm}, t = 1 \text{ min} = 60, T_1 - T_2 = 100^\circ\text{C}$$

$$Q = K \frac{A (T_1 - T_2)}{x} t = \frac{0.2 \times 200 \times 100 \times 60}{0.2} = 1.2 \times 10^6 \text{ cal}$$

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Example (10) a rod with a length of 30cm and a cross-sectional area equal to 5cm² consists of two parts of equal length. Part **AB** is made of copper and part **BC** is made of iron and their point of connection is point **B**. As shown in the figure, the end **A** is kept at a temperature of 200 °C and the end **C** at a temperature of 0 °C. If the sides of the rod are thermally isolated, calculate the rate of heat flow along the length of the rod when the steady state is reached. Given that $K_1 = K_{copp} = 0.09 \frac{\text{cal}}{\text{sec.cm.}^\circ\text{C}}$, $K_2 = K_{iron} = 0.12 \frac{\text{cal}}{\text{sec.cm.}^\circ\text{C}}$.

Solution:



$$x_1 = x_2 = 15 \text{ cm}, A = 5 \text{ cm}^2, \Delta T_1 = 200 - T, \Delta T_2 = T - 0$$

Since steady state has been reached, the rate of heat flow through first part AB is equal to the rate of heat flow through second part BC, so

$$\left(\frac{dQ}{dt}\right)_1 = \left(\frac{dQ}{dt}\right)_2 \rightarrow K_1 \frac{A (200 - T)}{x_1} t = K_2 \frac{A (T - 0)}{x_2} t$$

$$K_1 (200 - T) = K_2 (T - 0) \rightarrow 0.09(200 - T) = 0.12 T$$

$$18 - 0.09T = 0.12T \rightarrow 0.21T = 18 \rightarrow T = 85.71^\circ\text{C}$$

$$\left(\frac{dQ}{dt}\right)_1 = K_1 \frac{A (200 - T)}{x_1} = 0.09 \times \frac{5(200 - 85.71)}{15} = 3.43 \text{ cal/sec}$$

Example (11) a sheet of asbestos paper with a thickness of 2 mm is used as a barrier between two brass plates. The temperature on one plate is equal to 1000°C and the other plate has a temperature of 200°C. What is the amount of heat passing through an area of 50 cm² from one plate to the other during the time t=20 min)? Note that the thermal conductivity coefficient of asbestos is $k=5 \times 10^{-4} \text{ cal/cm sec } ^\circ\text{C}$.

The solution

$$Q = k A \frac{T_1 - T_2}{x} t = 5 \times 10^{-4} \times 50 \times \frac{1000 - 200}{0.2} \times 20 \times 60$$

$$Q = 1.2 \times 10^5 \text{ cal}$$

Convective heat transfer:

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When heat is transferred in a physical medium by the transfer of some of its parts, the heat is said to be transferred by convection. This is due to the fact that in this scenario, certain portions of the medium carry heat with them as they are being transported. This kind of heat transfer takes place in fluids (including liquids and gases), and we refer to the heat transfer as being caused by convection currents, which can be broken down into two distinct categories.

Free convection currents: It is a phenomenon resulting from the effect of gravity and this transfer process is governed by the following law:

$$Q = \alpha A h t (T_1 - T_2) ,$$

Where T_1 and T_2 are the temperatures of the hot and cold regions, respectively. In addition, α is the convection coefficient, h is the thickness of the hot medium, A area of the hot medium and t is the heat transfer time.

Forced convection currents:

In this type of conduction, the fluid is forcibly moved by fans or pumps.

Newton's Law of Cooling: Newton's Law of Cooling states, "The rate of heat loss from a hot object $\frac{\Delta Q}{\Delta t}$ to the surrounding medium is proportional to the difference between the temperature of the hot object T and the temperature of the medium T_r .

$$\frac{\Delta Q}{\Delta t} = K (T - T_r)$$

Example (12) If we make a flat vertical wall with an area of 6 m^2 at a constant temperature of 116°C while the air on its sides at a temperature of 35°C , what is the heat lost from the wall in an hour due to natural convection. Knowing that the wall thickness is 1 m and the convection coefficient in this case $\alpha = 1.27 \times 10^{-3} \text{ cal/m}^2\text{s } ^\circ\text{C}$

Assume that the amount of heat transferred from each side can be found from the relationship:

$$Q = \alpha A h t (T_1 - T_2) = 1.27 \times 10^{-3} \times 6 \times 1 \times 60 \times 60 \times (116 - 35) = 2222 \text{ cal}$$

Radiation heat transfer:

In this type of heat transfer method, heat is transferred without the need for a means of transportation as thermal radiation is electromagnetic waves and the relationship between the amount of heat transferred by radiation and temperature is described by Stefan-Boltzmann's

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law, which states that ((the amount of heat radiated by a black body per unit area per unit time is proportional to the fourth power of the absolute temperature)),

$$\frac{dQ}{dt} = \sigma A(\Delta T)^4$$

Black body: A black body is an object that absorbs all the radiation energy falling on it and radiates energy continuously, and the best example of a black body is a gap in a wall. When any radiation enters, it suffers from several reflections inside the gap, so it loses all its energy during successive reflections inside the gap, and this indicates that the black body's absorption coefficient is equal to one and therefore its reflection coefficient is equal to zero .

The radiative power of a black body:

It is known that hot objects lose heat at greater rates than cold objects. This phenomenon was demonstrated by the two scientists Steifen and Boltzmann after series of experiments in what is known as Stephen's law of radiation, which states that “the total radiation rate emitted from the unit area of the black body is proportional to the fourth exponent of its absolute temperature”.

That is, $E = \sigma T^4$, where σ is a constant known as Stephen's constant and is equal to $(5.672 \times 10^{-8} \text{ J/m}^2\text{sk})$. For any body (which is not quite black), another amount is added called the degree of blackness ϵ which gives the amount of blackness in the ratio between the radiation emitted by the object to the radiation emitted by the black body at the same temperature.

For a non-black body, we write the Stephen-Boltzmann law as follows: $E = \epsilon \sigma T^4$, so the amount of heat transferred by radiation from a hot object to the surrounding medium can be written as $E = A \epsilon \sigma (T^4 - T_1^4)$ where T and T_1 are the temperatures of the hot object and the ambient medium, respectively.

Absorptive capacity and Kirchhoff's law :

The radiation capacity (α) is the ratio between the amount of radiation absorbed by an object to the amount of radiation falling on it, and Kirchhoff's law expresses the relationship between radiation capacity and absorption capacity. This law states that the ratio between the radiant power of a natural object and its absorptive capacity is a constant amount.

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Example (13) a black spherical object has diameter of 10 cm is stationary at a constant temperature of 327°C and the temperature of the medium is 27°C. What is the power radiated?

Consider $\varepsilon=1$ knowing that $\sigma = 5.67 \times \frac{10^{-8}W}{m^2K^4}$.

The solution

$$\begin{aligned} E &= \varepsilon \sigma A (T^4 - T_1^4) = (5.67 \times 10^{-8})(4\pi(0.05)^2)\{ (327 + 273)^4 - (27 + 273)^4 \} \\ &= (5.67 \times 10^{-8})(4\pi(0.05)^2)\{ (600)^4 - (300)^4 \} = 216.43 \text{ W} \end{aligned}$$

Problems on chapter 1

Choose the one alternative that best completes the statement or answers the question:

1. In order to understand the concept of temperature it is necessary to understand
 - a. the zeroth law of thermodynamics.
 - b. the first law of thermodynamics.
 - c. the second law of thermodynamics.
 - d. all of the above.
 - e. only (b) and (c) above.
2. In order for two objects to have the same temperature, they must
 - a. be in thermal equilibrium.
 - b. be in thermal contact with each other.
 - c. have the same relative “hotness” or “coldness” when touched.
 - d. have all of the properties listed above.
 - e. have only properties (b) and (c) above.
3. A pressure of 10 mm Hg is measured at the triple-point of water using a constant-volume gas thermometer, what will the pressure be (in mm Hg) at 50°C?
 - a. 68.3
 - b. 1.8
 - c. 31.8
 - d. 11.8
 - e. 8.5
4. A pressure of 10 mm Hg is measured using a constant-volume gas thermometer at a temperature of 50°C. What is the pressure (in mm Hg) at the zero-point temperature?
 - a. 31.8
 - b. 11.8
 - c. 8.5
 - d. 54.6
 - e. 68.3
5. A temperature difference of 5 K is equal to
 - a. a difference of 9 on the Celsius scale.
 - b. a difference of 9 on the Fahrenheit scale.
 - c. a difference of 2.8 on the Rankine scale.
 - d. a difference of .5 on the Fahrenheit scale.
 - e. a difference of 2.8 on the Celsius scale.

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6. A thermometer registers a change in temperature of 100°F . What change in temperature does this correspond to on the Kelvin Scale?
- 453
 - 328
 - 180
 - 55.6
 - 24.5
7. Helium condenses into the liquid phase at approximately 4 K. What temperature, in degrees Fahrenheit, does this correspond to?
- -182
 - -269
 - -118
 - -452
 - -484
8. Two thermometers are calibrated, one in degrees Celsius and the other in degrees Fahrenheit. At what temperature (in kelvins) do their readings measure the same temperature?
- 218.15
 - 233.15
 - 273.15
 - 40.15
 - 0
9. A child has a temperature of 104°F . What is the temperature in degrees kelvin?
- 40
 - 406
 - 401
 - 313
 - 349
10. At what temperature is the Celsius scale reading equal to twice the Fahrenheit scale reading?
- -12.3°F
 - -24.6°F
 - -12.3°C
 - -6.1°C
 - -20°F

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11. A bridge is made with segments of concrete 50 m long. If the linear expansion coefficient is $12 \times 10^{-6} (\text{°C})^{-1}$, how much spacing (in cm) is needed to allow for expansion during an extreme temperature change of 150°F ?
- 10
 - 2.5
 - 7.5
 - 5.0
 - 9.5
12. A building made with a steel structure is 650 m high on a winter day when the temperature is 0°F . How much taller (in cm) is the building when it is 100°F ? (The linear expansion coefficient of steel is $11 \times 10^{-6} (\text{°C})^{-1}$.)
- 71
 - 36
 - 40
 - 46
 - 65
13. A gallon container is filled with gasoline. How many gallons are lost if the temperature increases by 25°F ? (The volume expansion of gasoline is $9.6 \times 10^{-4} (\text{°C})^{-1}$.) (Neglect the change in volume of the container.)
- 2.4×10^{-2}
 - 1.3×10^{-2}
 - 3.6×10^{-2}
 - 4.8×10^{-2}
 - 9.6×10^{-2}
14. An auditorium has dimensions $10 \text{ m} \times 10 \text{ m} \times 60 \text{ m}$. How many moles of air fill this volume at STP?
- 2.7×10^2
 - 2.7×10^4
 - 2.7×10^3
 - 2.7×10^5
 - 2.7×10^6

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15. An auditorium has a volume of $6 \times 10^3 \text{ m}^3$. How many molecules of air are needed to fill the auditorium at STP?
- a. 1.6×10^{29}
 - b. 1.6×10^{27}
 - c. 1.6×10^{25}
 - d. 1.6×10^{23}
 - e. 1.6×10^{20}
16. One mole of an ideal gas is held at a constant pressure of 1 atm. Find the change in volume (in liters) if the temperature changes by 50°C .
- a. 1
 - b. 2
 - c. 3
 - d. 4
 - e. 5
17. One mole of an ideal gas is held at a constant volume of 1 liter. Find the change in pressure if the temperature increases by 50°C .
- a. 3 atm
 - b. 4 atm
 - c. 2 atm
 - d. 1 atm
 - e. 5 atm
18. One mole of an ideal gas has a temperature of 25°C . If the volume is held constant and the pressure is doubled, the final temperature (in $^\circ\text{C}$) will be
- a. 174
 - b. 596
 - c. 50
 - d. 323
 - e. 25
19. A bicycle pump contains air at STP. As the tire is pumped up, the volume of air decreases by 50% with each stroke. What is the new pressure of air (in atm) in the chamber after the first stroke, assuming no temperature change?
- a. 2
 - b. 1
 - c. 0.5
 - d. 0.1
 - e. 3

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20. A helium-filled balloon has a volume of 1 m^3 . As it rises in the earth's atmosphere, its volume expands. What will its new volume be (in m^3) if its original temperature and pressure are 20°C and 1 atm , and its final temperature and pressure are -40°C and 0.1 atm ?
- 4
 - 6
 - 8
 - 10
 - 1.5
21. A bubble having a diameter of 1.00 cm is released from the bottom of a swimming pool where the depth is 5.00 m . What will the diameter of the bubble be when it reaches the surface? The temperature of the water at the surface is 20.0°C , whereas it is 15.0°C at the bottom. (The density of water is $1.00 \times 10^3 \text{ kg/m}^3$.)
- 1.05
 - 1.15
 - 1.45
 - 1.65
 - 1.35
22. A scuba diver has his lungs filled to half capacity (3 liters) when 10 m below the surface. If the diver holds his breath while quietly rising to the surface, what will the volume of the lungs be (in liters) at the surface? Assume the temperature is the same at all depths. (The density of water is $1.0 \times 10^3 \text{ kg/m}^3$.)
- 5.9
 - 4.5
 - 6.4
 - 3.9
 - 3.1
23. Two identical containers, A and B , hold equal amounts of the same ideal gas at the same P_0 , V_0 and T_0 . The pressure of A then decreases by a half while its volume doubles; the pressure of B doubles while its volume decreases by a half. Which statement correctly describes the temperatures of the gases after the changes?
- $T_A = 0.5T_B = T_0$.
 - $T_B = 0.5T_A = T_0$.
 - $T_B = T_A = T_0$.
 - $T_A = 2T_B = T_0$.
 - $T_B = 2T_A = T_0$.

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24. Which of the following is not a possible thermometric property of a body?
- The change in length of a solid.
 - the change in volume of a gas at constant pressure.
 - The change in pressure of a gas at constant volume.
 - The change in weight at constant pressure and volume.
 - The change in electrical resistance of a conductor.
25. A pebble size object and a bowling ball size probe from a spaceship land on a large asteroid that is far from any star. After a long period of time has passed, it is highly probable that the pebble and the probe
- have each had the same change in temperature.
 - have each had the same change in volume.
 - are in thermal equilibrium with one another.
 - are not in thermal equilibrium with one another.
 - are in thermal equilibrium with one another, but are not at the same temperature.
26. A temperature difference of 9 Celsius degrees is equal to a Fahrenheit temperature difference of
- 5 Fahrenheit degrees.
 - 9 Fahrenheit degrees.
 - 16 Fahrenheit degrees.
 - 37 Fahrenheit degrees.
 - 41 Fahrenheit degrees.
27. Death Valley in California receives many German tourists. When you convert a summer temperature reading of 130°F to the Celsius scale they use at home, you find that the Celsius temperature is
- 26°C.
 - 54°C.
 - 72°C.
 - 176°C
 - 327°C.
28. A beaker is filled to the 500 ml mark with alcohol. What increase in volume (in ml) the beaker contain when the temperature changes from 5°C to 30°C? (Neglect the expansion of the beaker, evaporation of alcohol and absorption of water vapor by alcohol.) $\beta_{\text{alcohol}} = 1.12 \times 10^{-4} / ^\circ\text{C}$
- 0.47
 - 0.93
 - 1.4
 - 1.7
 - 2.5

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29. What is the change in area (in cm^2) of a 60.0 cm by 150 cm automobile windshield when the temperature changes from 0°C to 36.0°C . The coefficient of linear expansion of glass is $9 \times 10^{-6}/^\circ\text{C}$.
- 1.62
 - 2.92
 - 3.24
 - 4.86
 - 5.83
30. A container with a one-liter capacity at 27°C is filled with helium to a pressure of 2 atm. ($1 \text{ atm} = 1.0 \times 10^5 \text{ N/m}^2$.) How many moles of helium does it hold?
- 0.040
 - 0.080
 - 0.45
 - 0.90
 - 1.0
31. Two bodies can be in thermal equilibrium with one another when they are at the same temperature even if they
- absorb different quantities of thermal energy from their surroundings in equal time intervals.
 - have different masses.
 - have different volumes.
 - have any of the properties listed above.
 - have any of the properties listed above and one of them is contact with a third body at a different temperature.
32. Angela claims that she wears a cylindrical-shaped hollow gold bracelet because it expands less than a solid one with a change in temperature. Clarissa claims that a cylindrical-shaped solid gold bracelet expands less than a hollow one. Which one, if either, is correct?
- Angela, because the bracelet expands outward on its outer surface and inward on its inner surface.
 - Clarissa, because the bracelet expands outward on its outer surface and inward on its inner surface.
 - Angela, because the inner circumference does not change, but the outer circumference expands.
 - Clarissa, because the inner circumference does not change, but the outer circumference expands.
 - Neither, because both the inner and outer circumferences increase in length.

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33. A student has written the equation below to convert a temperature in degrees Fahrenheit into Kelvins. What is wrong with this equation?

$$T_K = \frac{9}{5}T_F + 32$$

- The factor in front of T_F should be $\frac{5}{9}$.
 - The numerical factor $\frac{5}{9}$ should multiply $(T_F - 32)$.
 - An additional 273.15 Kelvins must be added to the right side of the equation.
 - All the corrections above are required.
 - Only corrections (b) and (c) are required.
34. Two moles of an ideal gas are placed in a container of adjustable volume. When measurements are made
- the pressure is inversely proportional to the volume at constant temperature.
 - the temperature is directly proportional to the volume at constant pressure.
 - the temperature is directly proportional to the pressure at constant volume.
 - all the statements above are found to be correct.
 - only statements (a) and (b) are found to be correct.
35. When the coefficient of linear expansion, α , and the temperature change, $T_f - T_i$, are large, a length L_i of a solid substance expands in length to
- $L_f = \alpha L_i (T_f - T_i)$.
 - $L_f = L_i [1 + \alpha (T_f - T_i)]$
 - $L_f = L_i [1 + \ln(\alpha (T_f - T_i))]$.
 - $L_f = L_i e^{\alpha (T_f - T_i)}$.
 - $L_f = L_i [1 + e^{\alpha (T_f - T_i)}]$.
36. A square plate has an area of 29.00 cm^2 at 20.0°C . It will be used in a low temperature experiment at $T = 10.0 \text{ K}$ where it must have an area of 28.00 cm^2 . What area must be removed from the plate at 20.0°C for it to have the correct area at 10.0 K ? (The coefficient of linear expansion is $10 \times 10^{-6} (\text{C}^\circ)^{-1}$.)
- 0.0793 cm^2
 - 0.159 cm^2
 - 0.238 cm^2
 - 0.841 cm^2
 - 0.921 cm^2

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37. Equal volumes of hydrogen and helium gas are at the same pressure. The gram molecular mass of helium is four times that of hydrogen. If the total mass of both gases is the same, the ratio of the temperature of helium (He) to that of hydrogen (H_2) is
- $\frac{1}{4}$.
 - $\frac{1}{2}$.
 - 1.
 - 2.
 - 4.
38. Equal masses of hydrogen and helium gas are at the same temperature in vessels of equal volume. The gram molecular mass of helium is four times that of hydrogen. If the total mass of both gases is the same, the ratio of the pressure of helium (He) to that of hydrogen (H_2) is
- $\frac{1}{4}$.
 - $\frac{1}{2}$.
 - 1.
 - 2.
 - 4.
39. Steel blocks A and B, which have equal masses, are at $T_A = 300^\circ\text{C}$ and $T_B = 400^\circ\text{C}$. Block C, with $m_C = 2m_A$, is at $T_C = 350^\circ\text{C}$. Blocks A and B are placed in contact, isolated, and allowed to come into equilibrium. Then they are placed in contact with block C. At that instant,
- $T_A = T_B < T_C$.
 - $T_A = T_B = T_C$.
 - $T_A = T_B > T_C$.
 - $T_A + T_B = T_C$.
 - $T_A - T_B = T_C$.

40-If a Celsius thermometer reads 30 degrees Celsius, what will a Fahrenheit thermometer read in the same room

- A) 303 K B) 86°C C) 86 f D) 48.66 f

41- Before ice can form on a lake, all the water in the lake must be cooled to

- A) 4°C . B) 0°C . C) A value slightly below 0°C . D) Neither of these

Ch1: Temperature and Thermal Energy

42- a rod made of silver is exactly 200 cm long at a temperature of 20°C. Calculate its length when the temperature is 220°C, given that the coefficient of longitudinal expansion of silver is $\alpha = 18 \times 10^{-6} \text{ } ^\circ\text{C}^{-1}$

- A) 200.72 cm B) 200.36 cm C) 99.28 cm D) None of these

43- The human body loses heat by

- A) radiation. B) convection. C) conduction. D) all of these

44- a black spherical object has area of 1.8 m^2 is stationary at a constant temperature of 227°C and the temperature of the medium is 27°C. What is the power radiated? Consider $\epsilon = 1$

knowing that $\sigma = 5.67 \times \frac{10^{-8} \text{ W}}{\text{m}^2 \text{ K}^4}$.

- A) 6379 watt B) 2701 watt C) 5552 watt D) 271 watt

45- Increasing the temperature of 40 grams of water by 1°C requires

- A) 4.186 calorie. B) 1 calories. C) 40 joul D) 167.44 joul

46 - a paper slide is 2 cm thick and has a surface area of 50 cm^2 . If the temperature difference between the two opposite surfaces is 100°C, find the amount of heat that will transferred through the slide in a time of one minute, given that $k = 2 \times 10^{-3} \text{ cal/sec.cm. } ^\circ\text{C}$

- A) 360 cal B) 720 cal C) 300 cal D) 600 cal

Heat and the First Law of Thermodynamics

Our focus in this chapter will be on the concept of internal energy, the first law of thermodynamics, and some applications of this law. The first law of thermodynamics expresses the general principle of conservation of energy. According to this law, an energy transfer to or from a system by either heat or work can change the internal energy of the system.

Heat and Thermal Energy

It is important to make a major distinction between *heat* and *internal energy (thermal energy)*.

Internal energy is all the energy of a system that is associated with its microscopic constituents. Internal energy includes kinetic energy of random translational, rotational, and vibrational motion of molecules, potential energy of molecules and between molecules.

Heat is defined as the *transfer of energy* from one system to another due to a temperature difference between them.

Units of Heat, The Mechanical Equivalent of Heat

Previously, heat was measured in terms of its ability to raise the temperature of water. Thus, the **calorie** (cal), in cgs units, was defined as the amount of heat required to raise the temperature of 1 g of water from 14.5 to 15.5 °C. The ‘Calorie’ with a capital C, used by nutritionists, is a kilocalorie ($1\text{Cal}=1\text{ kcal}=10^3\text{ cal}$).

The **British Thermal Unit** (Btu) was also defined as the amount of heat required to raise the temperature of 1 lb of water from 63 to 64 °F. Since heat is now known as transferred energy, the SI unit for it is the **joule** (J).

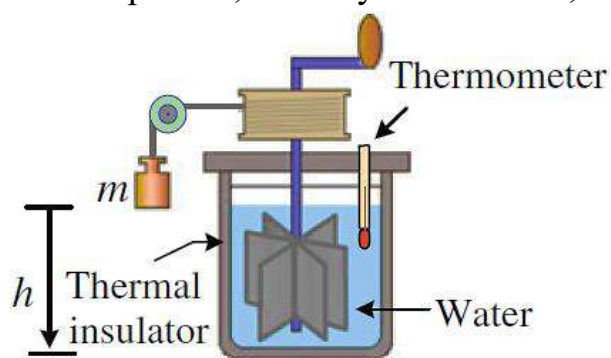


Fig.1 Joule's experiment for measuring the mechanical equivalent of heat from the temperature rise in water

In a famous experiment, see Fig. 1, Joule measured the calorie (cal) by converting mechanical energy into heat energy, expressed as an increase in water temperature. Joule found that the loss in mechanical energy is proportional to the increase in temperature of the water. The proportionality constant was found to be equal to 4,180 J/kg. °C. Hence, 4,180 J of mechanical

energy will raise the temperature of 1 kg of water from 14.5 to 15.5 °C. One kilocalorie (1 kcal) is now defined to be exactly 4,186 J without reference to the heating of substance. Thus:

$$1 \text{ kcal} = 4,186 \text{ J.}$$

Heat and Work

In thermodynamics, when an isolated system is in thermal equilibrium internally, we describe its macroscopic state with the variables P , V , and T to represent pressure, volume and temperature. For such a system, we describe its microscopic state of internal energy with the variable E_{int} (some other textbooks use the symbol U).

Let us assume our system consists of gas confined to a cylinder with insulated walls and a movable frictionless piston of area A , as shown in Fig. 2. The cylinder rests on a heat reservoir whose temperature T is controlled by a knob. At equilibrium, the upward force on the piston due to the pressure of the confined gas is equal to the weight of the load on the top of the piston.

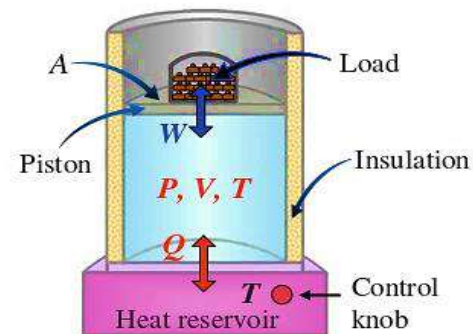


Fig.2 Gas confined to a cylinder with a movable frictionless piston. The gas can do work W to raise or lower the piston. By regulating the temperature T of the thermal reservoir, by means of a control knob, a quantity of heat Q can be added or removed from the gas

Consider that we start at an *initial state* i , where the system is described to have pressure P_i , volume V_i , and temperature T_i . We then change the system to a *final state* f , described to have pressure P_f , volume V_f , and temperature T_f . The process of changing the system from the initial state to the final state is a **thermodynamic process**. During such a process, work is done *by* the system to raise the piston (positive work) or lower it (negative work). In addition, heat may be transferred into the system from the thermal reservoir (positive heat) or vice versa. We assume that the state of the gas changes **quasi-statically**, i.e. slowly enough to allow the system to remain essentially in a thermodynamic equilibrium at all times.

Now, assume we reduce the load from the piston in such a way that the piston will move upward through a differential displacement $\vec{ds}$ with *almost* constant upward force $\vec{F}$, as shown in Fig. 3. From the definition of pressure, we have $F = PA$, where A is the area of the piston. The differential

work dW done by the gas during the displacement is: $dW = \vec{F} \cdot \vec{ds} = F ds = P A ds$, Since $A ds$ is the differential change in the volume of the gas dV (i.e. $dV = A ds$), we can express the work done by the gas as follows: $dW = P dV$

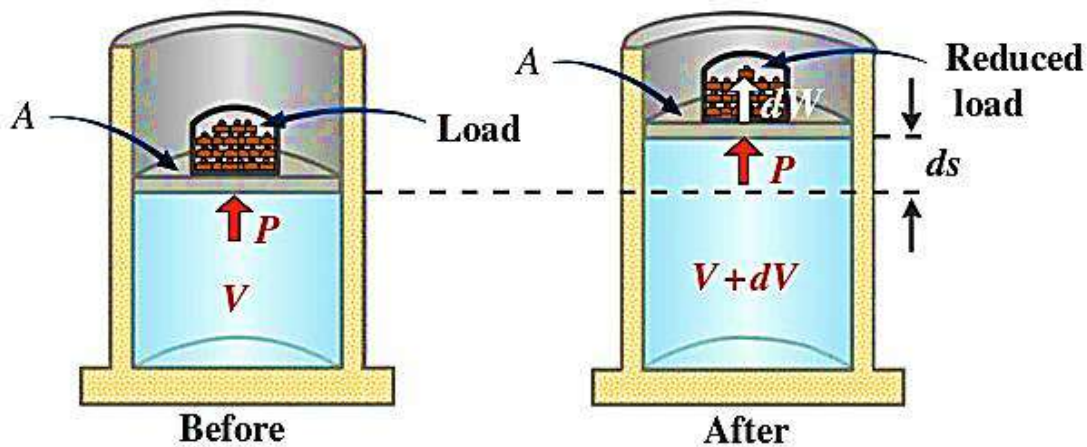


Fig. 3 A confined gas in a cylinder at pressure P does work dW on a free piston as the gas expands from volume V to volume $V + dV$ because of a decreased load

If the gas expands, as in Fig. 4, then dV is positive and the work done by the gas is positive, whereas if the gas is compressed, dV is negative, indicating that the work done by the gas is negative (which can be interpreted as work done *on* the gas). When we remove an appreciable amount of load from the piston, the volume of the gas changes from V_i to V_f , and the total work done by the gas is:

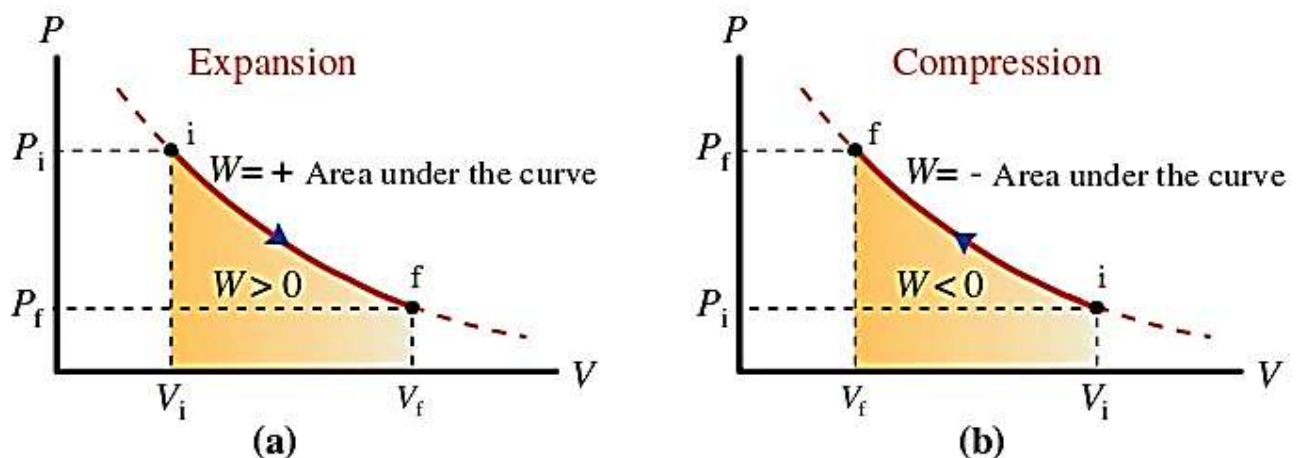
$$W = \int_{V_i}^{V_f} dw = \int_{V_i}^{V_f} P dV$$


Fig. 4 The figure shows a gas that goes from an initial state i to a final state f by means of a thermodynamic process. (a) When the gas expands, the work done by the gas is positive and equals the area under the PV curve. (b) Similar to (a), except that the gas is compressed and the work done by the gas is negative

Chapter 2 Heat and the First Law of Thermodynamics

As seen from Fig. 4, the total work done during the expansion or compression of the gas depends on the specific path taken from the initial state i to the final state f . In Fig. 5, we illustrate this important point further by considering several different paths for the gas along the PV curve, from state i to state f , regardless of how we achieve each path.

Path a: The gas expands from V_i to V_f while the pressure decreases from P_i to P_f . The work done by the gas along this path is positive and represented by the colored area under the curve between i and f .

Path b: The gas first expands from V_i to V_f at constant pressure P_i , and then its pressure is reduced to P_f at constant volume V_f . The work done along this path is $P_i (V_f - V_i)$.

Path c: The pressure of the gas is first reduced from P_i to P_f by cooling at a constant volume V_i , and then allowing the gas to expand from V_i to V_f at constant pressure P_f . The work done along this path is $P_f (V_f - V_i)$.

Path d: The gas is compressed from V_i to V_f while the pressure increases from P_i to P_f . The work done along this path is the negative of the colored area under the curve.

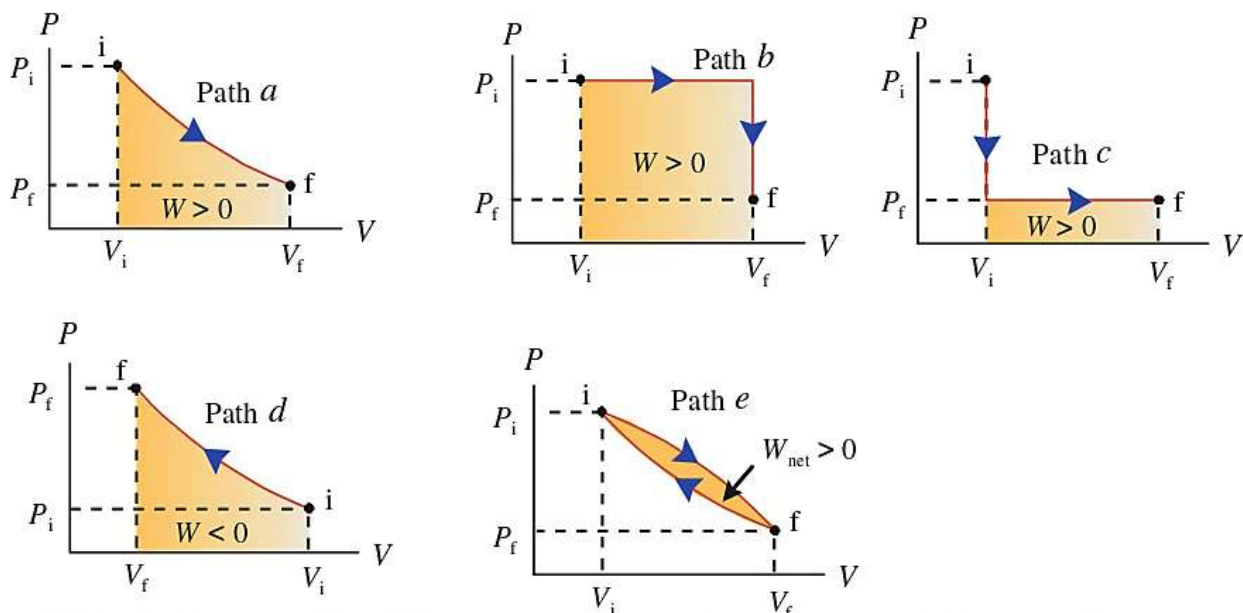


Fig.5 The gas of Fig. 2 goes from an initial state i to a final state f by means of several different thermodynamic processes

Path e: The net work done by the system during a closed cycle is the sum of the positive work done during the expansion and the negative work done during the compression. Here, the net

work done by the gas is positive and is represented by the enclosed area between the two curves. From the graphs of Fig. 5, we see that W could be small or large depending on the thermodynamic path between i and f. Thus:

Spotlight

The net work done by a system W depends on the thermodynamic process (or the path) chosen between its initial and final states.

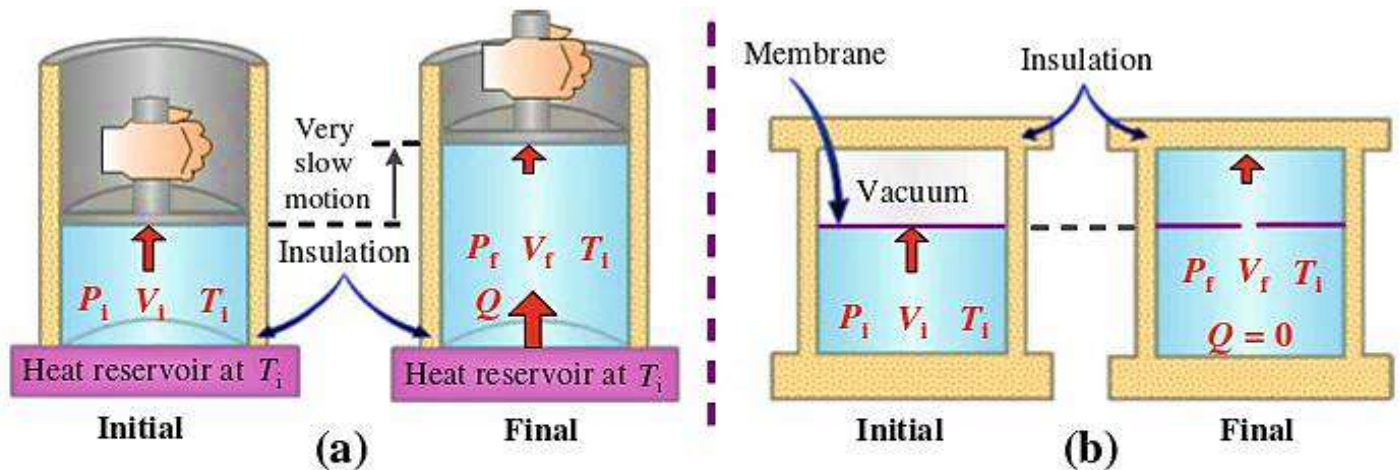


Fig. 6 (a) An ideal gas at temperature T_i expands slowly while absorbing heat energy Q from a reservoir in order to maintain its constant temperature T_i . (b) An ideal gas expands rapidly into an evacuated chamber after a membrane is broken

In a similar manner, we also find that the heat energy transfer Q into or out of a system depends on the thermodynamic process. This can be demonstrated for an ideal gas as shown in Fig. 6.

In Fig. 6 a, the piston is held at a position where the gas is at its initial pressure P_i , volume V_i , and temperature T_i . When the force holding the piston is reduced slightly, the piston rises very slowly to a final pressure P_f and final volume V_f , i.e. the gas is doing work W on the piston. During this expansion process, heat energy Q is transferred from the reservoir to the gas to maintain a constant temperature T_i .

In Fig. 6 b, the thermally insulated gas has the same initial state as in Fig. 6a, but with a membrane replacing the piston. When the membrane is broken, the gas expands *rapidly* into the vacuum until it acquires a pressure P_f and volume V_f . In this case, the gas does no work, i.e.

$W=0$, and no heat is transferred, i.e. $Q = 0$.

In both parts of Fig. 6, the initial and final states of the ideal gas are identical, although the path is different. In part (a) of the figure the gas does work W on the piston, and heat energy Q is transferred slowly to the gas from the reservoir. In part (b) of the figure the work done by the gas is zero and no heat energy is transferred. Thus:

Spotlight

The heat energy transfer Q *depends on the thermodynamic process (or the path) chosen between the initial and final states of a system.*

Finally, we conclude that neither the work done nor the heat energy are independently conserved during a thermodynamic process between the initial and final states of a system.

The First Law of Thermodynamics

The principle of conservation of energy as applied to systems that are *not isolated*, can be expressed as $W = E_{\text{tot}} = \Delta K + \Delta U + \Delta E_{\text{int}}$. In this chapter, we assume that there are no changes in kinetic energy and potential energy of the system as a whole; that is, $\Delta K = \Delta U = 0$ and hence

$W = \Delta E_{\text{tot}} = \Delta E_{\text{int}}$. Moreover, before this chapter, the term *work* and the symbol W always meant the work done *on* a system. But starting from equation ($dW = P dV$), and continuing to the rest of this chapter, we focus on the work done *by* a system. Thus, we replace the symbol W by $-W$ and hence becomes $-W = \Delta E_{\text{tot}} = \Delta E_{\text{int}}$. If we need to account for the transfer of heat energy Q that is added (if Q positive) or taken (if Q negative) from the system, then we add Q to the left hand side of this equation and arrive at the following thermodynamic equation: $\Delta E_{\text{int}} = Q - W$ (The first law of thermodynamics). As we saw, W and Q are path-dependent, yet a surprising experimental discovery was found: *The quantity $Q - W$ is the same for all thermodynamic processes.* It depends only on the initial and final states of the system and is *path-independent*.

Equation $\Delta E_{\text{int}} = Q - W$ is known as the **first law of thermodynamics**. This law states that a change in internal energy in a system can occur as a result of energy transfer by heat or by work, or by both. If the thermodynamic system undergoes only a differential change, we can write the

first law as: $dE_{\text{int}} = dQ - dW$ (The first law of thermodynamics)

Spotlight: The internal energy E_{int} of a system increases if energy is added via heat Q and decreases if energy is lost via work W done *by* the system.

Some special cases of the first law of thermodynamics are as follows

1. Isolated Systems

Consider a system that is not interacting with its surroundings. In this case, no energy transfer by heat takes place, i.e. $Q = 0$, and the value of the work done by the system is zero, i.e. $W = 0$. Then, from the first law we have $\Delta E_{\text{int}} = 0$. Thus, we conclude that the internal energy of an isolated system remains constant.

$$E_{\text{int}} = \text{constant} \quad (\text{Isolated system})$$

2. Cyclic Processes

Consider a non-isolated system that is taken through a cyclic process, i.e. a process that starts and ends at the same state. In this case, the change in the internal energy must again be zero, i.e. $\Delta E_{\text{int}} = 0$. Then, from the first law we have:

$$\Delta E_{\text{int}} = 0 \quad \text{and} \quad Q = W \quad (\text{Cyclic process})$$

On the PV curve, a cyclic process appears as a closed curve as shown in path (e) of Fig. 5. For this clockwise cyclic path, the net work done by the system (and Q) equals the area enclosed by the path.

Applications of the First Law of Thermodynamics

The first law of thermodynamics relates the changes in internal energy of a system to transfers of energy by work W or heat Q , or both. In this section, we consider applications of the first law in processes in which certain restrictions are imposed.

1. Adiabatic Process

An adiabatic process is one that occurs so *rapidly* or occurs in *thermally insulated systems* during which no transfer of heat energy enters or leaves the system, i.e. $Q = 0$. With this restriction and the application of the first law of thermodynamics to an adiabatic process, we get:

$$Q = 0 \quad \text{and} \quad \Delta E_{\text{int}} = -W \quad (\text{Adiabatic process})$$

Figure 7 shows an idealized adiabatic process. Heat cannot enter or leave the system because of the insulation. The only way of transferring energy to the system is by work. We see in this figure that if a gas is compressed adiabatically such that W is negative, then ΔE_{int} is positive and hence the temperature of the gas increases. Conversely, if a gas expands adiabatically such that W is positive, then ΔE_{int} is negative, and hence the temperature of the gas decreases.

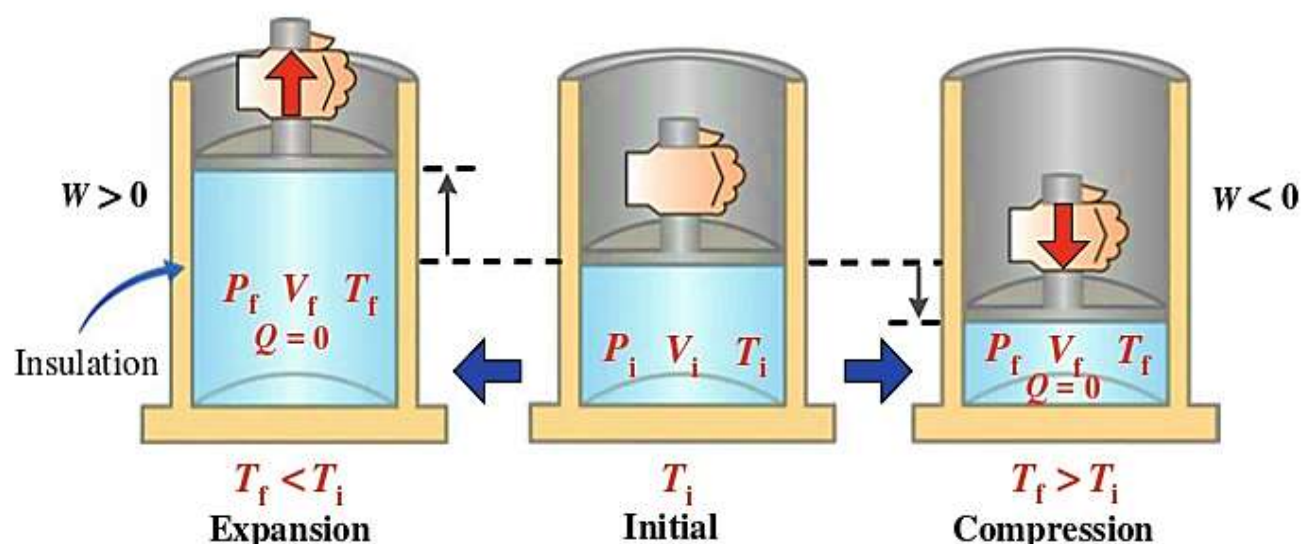


Fig. 7 An adiabatic compression/expansion is carried out for an ideal gas leading to an increase/decrease in internal energy

Adiabatic processes have a very important role in mechanical engineering. Some of the common examples include the approximately adiabatic compression/expansion of a mixture of gasoline vapor and air that takes place during operation of a combustion engine, leading to a temperature increase/decrease.

2. Adiabatic Free Expansion Process

The free expansion process is an adiabatic process, i.e. $Q = 0$, in which no work is done on or by the system, i.e. $W = 0$. Thus, with these restrictions and the application of the first law we have:

$$Q = W = 0 \quad \text{and} \quad \Delta E_{\text{int}} = 0 \quad (\text{Free expansion})$$

Figure 8 shows how such an expansion can be carried out. An ideal gas in thermal equilibrium is initially confined by a closed valve to one-half of an insulated chamber; the other half is evacuated. When we open the valve, the gas expands freely to fill both halves of the chamber. No heat is transferred to or from the gas because of the insulation. No work is done by the gas because it rushes into vacuum, during which its motion is unopposed by any counteracting pressure.

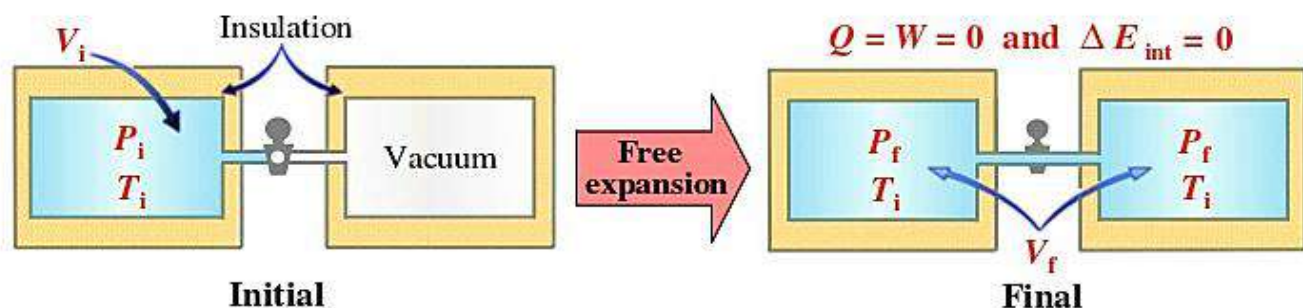


Fig. 8 In a free expansion process there will be no change in internal energy or temperature between the initial and final states

A free expansion differs from any other thermodynamic process since it cannot be performed slowly in a controlled way. As a result, at any given instant *during* the sudden expansion, the gas is not in thermal equilibrium and its pressure is not the same everywhere.

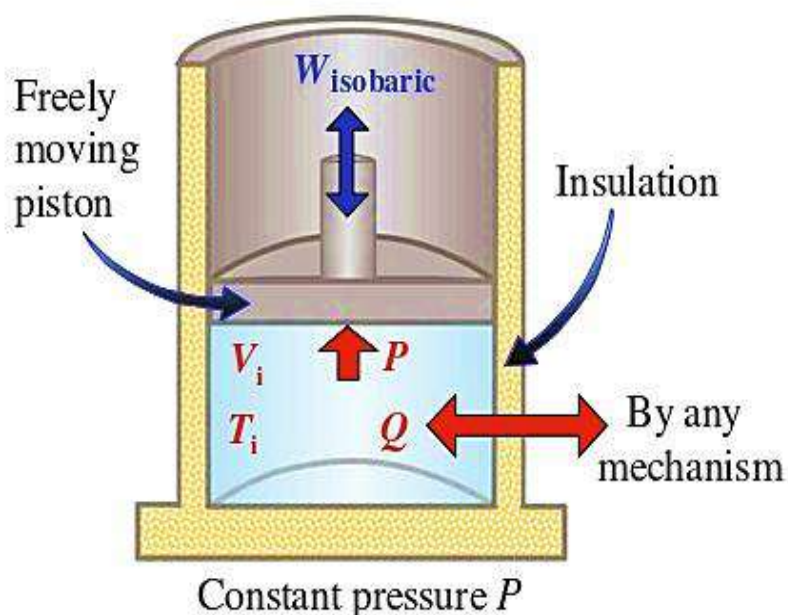
3. Isobaric Process

An isobaric process is one that takes place at constant pressure. In general, the first law of thermodynamics does not assume any special values for the isobaric process; that is, Q , W , and ΔE_{int} are all non-zero.

Assume the piston of Fig. 9 is free to move in such a way that it is always in equilibrium under the effect of the net force from a gas pushing upwards and

the weight of the piston plus the force due to atmospheric pressure pushing downwards. Then, an isobaric process could be established by transferring heat energy Q to or from the gas by any mechanism. This transfer causes the gas to expand or contract depending on the sign of Q . In the PV diagram of Fig. 5 the first process in path (b) and the second process in path (c) are examples of isobaric processes.

Fig. 9 An isobaric process could be achieved by transferring heat energy to a gas enclosed by a freely moving piston to attain a constant pressure



The work done by the gas as it expands or contracts in this isobaric process could be obtained from $W = \int P dV$, after removing the constant pressure from the integral, as follows:

$$W_{\text{isobaric}} = P(V_f - V_i) \quad (\text{Isobaric process})$$

4. Isovolumetric Process

An isovolumetric process is one that takes place at constant volume. In the PV diagram of Fig. 5, the second process in path (b) and the first process in path (c) are examples of isovolumetric processes.

Assume the piston of Fig. 10 is clamped to a fixed position to ensure an isovolumetric process. In such a process, the value of the work done by the gas is zero, i.e. $W = 0$, because the volume does not change. Thus, with this restriction and the application of the first law of thermodynamics to an isovolumetric process, we get:

$$W = 0 \quad \text{and} \quad \Delta E_{\text{int}} = Q \quad (\text{Isovolumetric process})$$

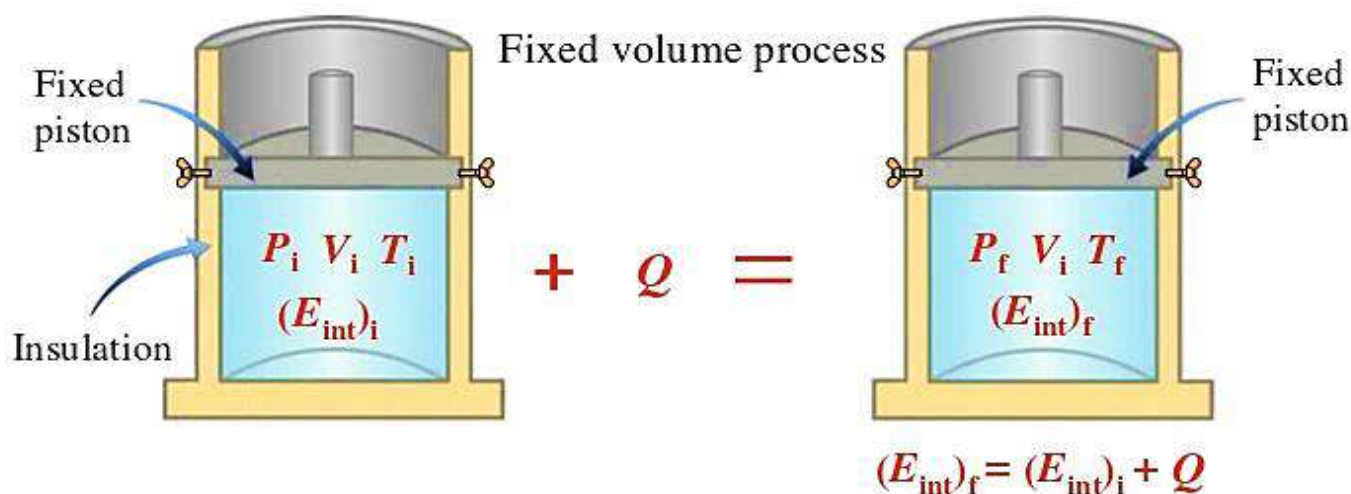


Fig. 10 An isovolumetric process could be achieved by fixing the piston's position. The pressure increases, and all the transferred heat energy remains in the system as an increase in its internal energy

This expression specifies that if energy is added by heat to a system kept at constant volume, then all of the transferred energy remains in the system as an increase in its internal energy, and hence, temperature. For example, when a closed metallic can is thrown into a fire, energy enters the gas in the can by the conduction of heat through the metal walls of the can. The temperature, and thus the pressure, in the can increases until the can possibly explodes, hence the warning label on such cans.

5. Isothermal Process

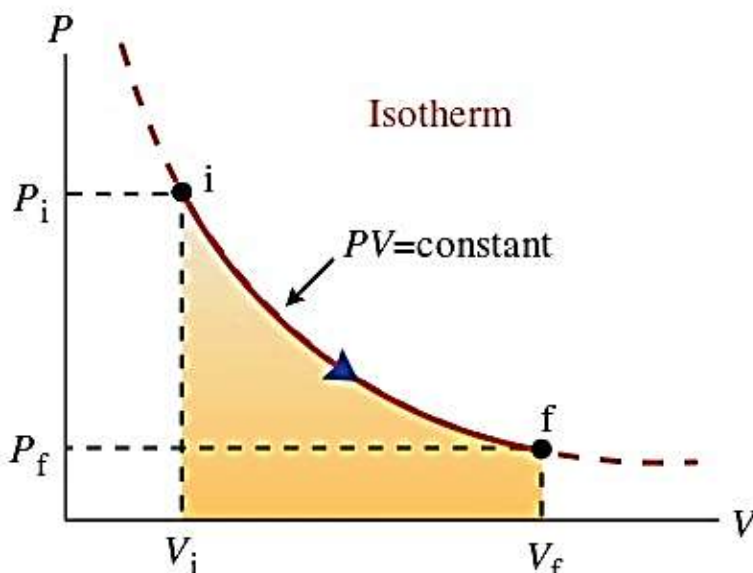
An isothermal process is one that takes place at constant temperature. This process can be established by putting a gas container in contact with a constant-temperature reservoir. If we plot P versus V at constant temperature for an ideal gas described by $\mathbf{P V = n R T}$, the plot yields a hyperbolic curve called an *isotherm*. Later we will prove that the internal energy of an ideal gas is a function of temperature only. Consequently, in an isothermal process involving an ideal gas we must have $\Delta E_{\text{int}} = 0$. Therefore, for an isothermal process, we conclude from the first law that the energy transfer Q must be equal to the work done by the gas W . That is:

$$\Delta E_{\text{int}} = 0 \quad \text{and} \quad Q = W \quad (\text{Isothermal process})$$

Any energy that enters the system by heat is transferred out of the system by work; as a result, no change in the internal energy of the system occurs in an isothermal process.

Suppose that an ideal gas is allowed to expand at constant temperature as described by the PV diagram in Fig. 11. According to $PV = NKT$, the curve is a hyperbola with the equation $PV = \text{constant}$.

Fig. 11 The PV diagram for an isothermal expansion of an ideal gas from initial state i to a final state f



Let us calculate the work done by the gas in the isothermal expansion from state i to state f , as shown in Fig. 11. Because the gas is ideal and the process is quasi-static, we can use the expression $PV = nRT$ for each point on the path. Therefore, we have:

$$W = \int_{V_i}^{V_f} P dV = \int_{V_i}^{V_f} \frac{nRT}{V} dV$$

Since T is constant and also n and R are constants, then they can be moved from the integral sign. Thus:

$$W = nRT \int_{V_i}^{V_f} \frac{dV}{V} = nRT \left| \ln V \right|_{V_i}^{V_f}$$

where we used $\int dV/V = \ln V$ to evaluate the last integral. Thus:

$$W = nRT \ln \left(\frac{V_f}{V_i} \right)$$

If the gas expands, the work W equals the positive of the shaded area under the PV curve shown in Fig. 11, this is because $\ln(V_f/V_i) > 0$. If the gas is compressed, $V_f < V_i$, then $\ln(V_f/V_i) < 0$ and the work done is the negative of the area under the PV curve.

Table summarizes the characteristics of the previous processes.

Table The first law of thermodynamics in five special cases

Process	Restriction	Consequence
Adiabatic	$Q = 0$	$\Delta E_{\text{int}} = -W$
Free expansion	$Q = W = 0$	$\Delta E_{\text{int}} = 0$
Isobaric	$P = \text{constant}$	$W_{\text{isobaric}} = P(V_f - V_i)$
Isovolumetric	$V = \text{constant}, W = 0$	$\Delta E_{\text{int}} = Q$
Isothermal (ideal gas)	$T = \text{constant}, \Delta E_{\text{int}} = 0$	$Q = W = nRT \ln(V_f/V_i)$

Example 1

At a constant pressure of 1 atm and a temperature of 0°C , the heat fusion of ice is $L_F = 3.33 \times 10^5 \text{ J/kg}$, the density of ice is $\rho_i = 920 \text{ kg/m}^3$, and the density of liquid water is $\rho_w = 1,000 \text{ kg/m}^3$. (a) Find the work W done by 1 kg of ice that melts completely to water. (b) Find the change in internal energy of this process.

Solution: (a) The initial volume of ice is:

$$V_i = m/\rho_i = (1 \text{ kg})/(920 \text{ kg/m}^3) = 1.087 \times 10^{-3} \text{ m}^3$$

The final volume of ice after it melts completely to water is:

$$V_w = m/\rho_w = (1 \text{ kg})/(1,000 \text{ kg/m}^3) = 10^{-3} \text{ m}^3$$

The work done by 1 kg of ice that melts completely to water under constant pressure of 1 atm ($1.01 \times 10^5 \text{ Pa}$) and temperature of 0°C , is:

$$\begin{aligned} W &= \int_{V_i}^{V_w} P dV = P(V_w - V_i) \\ &= (1.01 \times 10^5 \text{ Pa})(10^{-3} \text{ m}^3 - 1.087 \times 10^{-3} \text{ m}^3) = -8.787 \text{ J} \simeq -8.8 \text{ J} \end{aligned}$$

The minus sign appears because ice contracts when it melts.

(b) The heat energy transferred to change the phase of 1 kg of ice to water is:

$$Q = m L_F = (1 \text{ kg})(3.33 \times 10^5 \text{ J/kg}) = 3.33 \times 10^5 \text{ J}$$

Thus, from the first law of thermodynamics, we can find the change in internal energy of this process as follows:

$$\Delta E_{\text{int}} = Q - W = 3.33 \times 10^5 \text{ J} + 8.8 \text{ J} = 3.330088 \times 10^5 \text{ J}$$

We see from parts (a) and (b) that $|W|$ is less than 0.003% of Q in this process, i.e. $|W| \ll Q$. That is, the mechanical energy is negligible in comparison to the heat of fusion. So, all the added heat of fusion shows up as an increase in the internal energy.

Example 2

At a constant pressure of 1 atm, a movable piston encloses 1 kg of water with a volume of 10^{-3} m^3 and a temperature of 100°C , see Fig. 12. Heat is added from a reservoir until the liquid water changes completely into steam of volume 1.671 m^3 , see the figure. (a) How much work is done by the system (water + steam) during the boiling process? (b) How much heat energy is added to the system? (c) What is the change in the internal energy of the system?

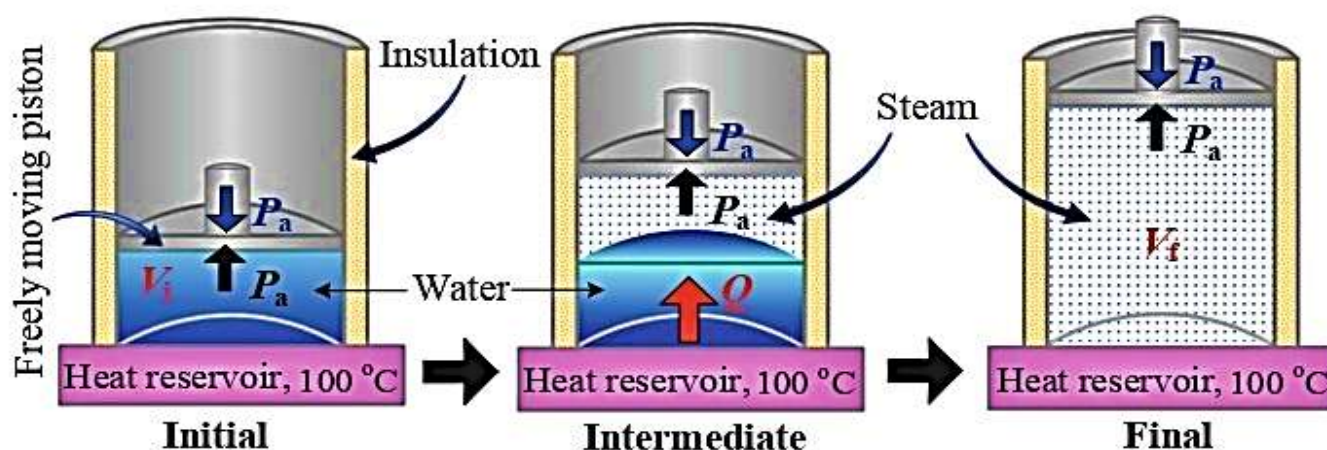


Fig. 12

Solution: (a) The work done by 1 kg of water that is converted completely into steam under a constant pressure of 1 atm ($1.01 \times 10^5 \text{ Pa}$) and a constant temperature of 100°C , is:

$$W = \int_{V_i}^{V_f} P dV = P(V_f - V_i) = (1.01 \times 10^5 \text{ Pa})(1.671 \text{ m}^3 - 10^{-3} \text{ m}^3) = 169 \text{ kJ}$$

(b) Since the heat of vaporization of water at atmospheric pressure is $2.26 \times 10^6 \text{ J/kg}$, the heat energy required to change the phase of 1 kg of water to steam will be:

$$Q = m L_V = (1 \text{ kg})(2.26 \times 10^6 \text{ J/kg}) = 2,260 \text{ kJ}$$

(c) From the first law of thermodynamics, we can find the change in internal energy of this process as follows:

$$\Delta E_{\text{int}} = Q - W = 2.26 \times 10^6 \text{ J} - 1.69 \times 10^5 \text{ J} = 2,091 \text{ kJ}$$

We see that about 92.5% of the heat energy goes into internal energy while the remaining 7.5% goes into external work.

Example 3

An aluminum rod of mass 1 kg is heated from 25 to 55 °C at constant atmospheric pressure, see Fig. 13. The aluminum rod has a density ρ of $2.7 \times 10^3 \text{ kg/m}^3$, a coefficient of volume expansion β of $7.2 \times 10^{-5} (\text{C}^\circ)^{-1}$ and a specific heat c of 900 J/kg.C° . (a) How much work is done by the rod? (b) How much heat is transferred to the rod? (c) Quantify the rod's internal energy change.

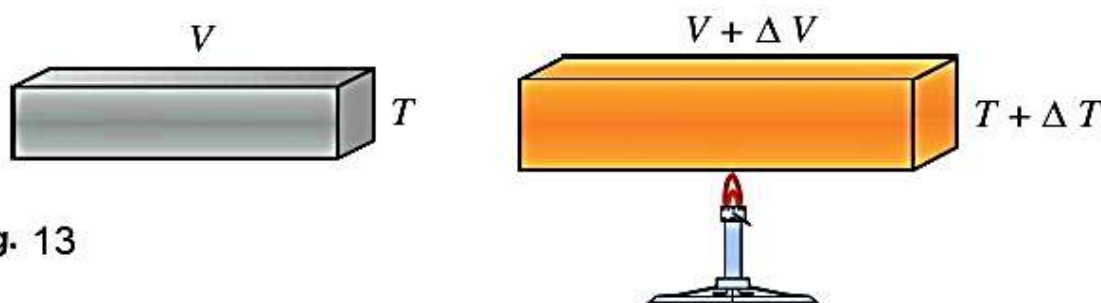


Fig. 13

Solution: (a) The initial volume of the aluminum rod is given by:

$$V = \frac{m}{\rho} = \frac{1 \text{ kg}}{2.7 \times 10^3 \text{ kg/m}^3} = 3.704 \times 10^{-4} \text{ m}^3$$

Using the change in temperature $\Delta T = T_f - T_i = 55^\circ\text{C} - 25^\circ\text{C} = 30^\circ\text{C}$, the change in the rod's volume can be obtained from Eq. 14.1 as follows:

$$\begin{aligned}\Delta V &= \beta V \Delta T \\ &= (7.2 \times 10^{-5} \text{ (C}^\circ\text{)}^{-1})(3.704 \times 10^{-4} \text{ m}^3)(30 \text{ C}^\circ) = 8 \times 10^{-7} \text{ m}^3\end{aligned}$$

Since the expansion is carried out at a constant pressure, the work done by the aluminum rod is:

$$\begin{aligned}W &= \int_{V_i}^{V_f} P dV = P(V_f - V_i) = P\Delta V \\ &= (1.01 \times 10^5 \text{ Pa})(8 \times 10^{-7} \text{ m}^3) = 8.08 \times 10^{-2} \text{ J}\end{aligned}$$

(b) We use the specific heat value in Eq. to calculate the amount of heat transferred to the rod as follows:

$$Q = mc \Delta T = (1 \text{ kg})(900 \text{ J/kg}\cdot\text{C}^\circ)(30 \text{ C}^\circ) = 2.7 \times 10^4 \text{ J}$$

(c) From the first law of thermodynamics, we can find the change in internal energy of this process as follows:

$$\Delta E_{\text{int}} = Q - W = 2.7 \times 10^4 \text{ J} - 8.09 \times 10^{-2} \text{ J} = 2.699 \times 10^4 \text{ J}$$

We notice that almost all of the heat energy goes towards increasing the internal energy of the aluminum rod. The fraction of heat energy that is used as work against the atmospheric pressure is only about $4 \times 10^{-4}\%$. Therefore, in thermal expansion of solids, the amount of energy that goes into work is usually neglected.

Example 4

Find the work done by 1 kmol of an ideal gas that is kept at a constant temperature of 27°C in an expansion process from 2 to 5 L.

Solution: Rewriting these values and the gas constant R , we have:

$$n = 1 \text{ kmol}$$

$$R = 8.314 \times 10^3 \text{ J/kmol}\cdot\text{K}$$

$$T = 27^\circ\text{C} = 27 + 273 = 300 \text{ K}$$

$$V_i = 2 \text{ L} = 2,000 \text{ cm}^3 = 2 \times 10^{-3} \text{ m}^3$$

$$V_f = 5 \text{ L} = 5,000 \text{ cm}^3 = 5 \times 10^{-3} \text{ m}^3$$

Since this process is isothermal, the work done by the ideal gas is given by

$$\begin{aligned} W &= nRT \ln \left(\frac{V_f}{V_i} \right) \\ &= (1 \text{ kmol})(8.314 \times 10^3 \text{ J/kmol.K})(300 \text{ K}) \ln \left(\frac{5}{2} \right) = 2.29 \times 10^6 \text{ J} \end{aligned}$$

This means that the heat energy Q that must be given to the ideal gas from the reservoir to keep its temperature $T = 27^\circ\text{C}$ is also $2.29 \times 10^6 \text{ J}$.

Answer the following questions

Multiple Choice

1. The work done in the expansion from an initial to a final state
 - a. is the area under the curve of a PV diagram.
 - b. depends only on the end point.
 - c. is independent of the path.
 - d. is the slope of a PV curve.
 - e. equals $P(V_F - V_i)$.
2. Gas in a container expands at a constant pressure of 3 atm. Find the work done (in J) by the gas if the initial volume is 5 liters and the final volume is 10 liters.
 - a. 0
 - b. 150
 - c. 15
 - d. 1500
 - e. 1.5
3. Gas in a container increases its pressure from 1 atm to 3 atm while keeping its volume constant. Find the work done (in J) by the gas if the volume is 5 liters.
 - a. 0
 - b. 3
 - c. 5
 - d. 15
 - e. 15×10^2
4. In an adiabatic free expansion
 - a. no heat is transferred between a system and its surroundings.
 - b. the pressure remains constant.
 - c. the temperature remains constant.
 - d. the volume remains constant.
 - e. the process is reversible.
5. In an isothermal process
 - a. the volume remains constant.
 - b. the temperature remains constant.
 - c. no heat is transferred between a system and its surroundings.
 - d. the pressure remains constant.
 - e. the internal energy is not constant.

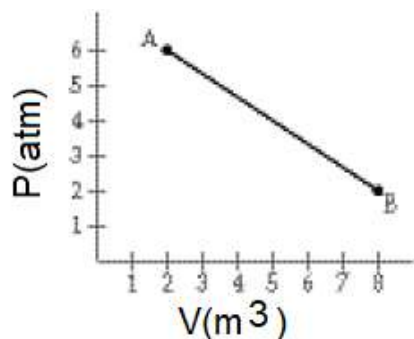
6. In an isobaric process
- the volume remains constant.
 - the temperature remains constant.
 - the pressure remains constant.
 - no heat is transferred between a system and its surroundings.
 - the internal energy is constant.
7. In an isovolumetric process
- the temperature remains constant.
 - no heat is transferred between a system and its surroundings.
 - the pressure remains constant.
 - the volume remains constant.
 - the internal energy is constant.
8. Determine the work done by 5 moles of an ideal gas that is kept at 100°C in an expansion from 1 liter to 5 liters.
- $2.5 \times 10^4 \text{ J}$
 - $1.1 \times 10^4 \text{ J}$
 - $6.7 \times 10^3 \text{ J}$
 - $2.9 \times 10^3 \text{ J}$
 - $8.4 \times 10^3 \text{ J}$
9. One gram of water is heated from 0°C to 100°C at a constant pressure of 1 atm. Determine the approximate change in internal energy (in cal) of the water.
- 160
 - 130
 - 100
 - 180
 - 50
10. Five moles of an ideal gas expands isothermally at 100°C to five times its initial volume. Find the heat flow into the system.
- $2.5 \times 10^4 \text{ J}$
 - $1.1 \times 10^4 \text{ J}$
 - $6.7 \times 10^3 \text{ J}$
 - $2.9 \times 10^3 \text{ J}$
 - $7.0 \times 10^2 \text{ J}$

11. An 8,000-kg aluminum flagpole 100-m long is heated by the sun from a temperature of 10°C to 20°C . Find the work done (in J) by the aluminum if the linear expansion coefficient is $24 \times 10^{-6} (\text{^{\circ}\text{C}})^{-1}$. (The density of aluminum is $2.7 \times 10^3 \text{ kg/m}^3$ and $1 \text{ atm} = 1.0 \times 10^5 \text{ N/m}^2$.)
- 287
 - 425
 - 213
 - 710
 - 626
12. An 8 000-kg aluminum flagpole 100-m long is heated by the sun from a temperature of 10°C to 20°C . Find the heat transferred (in J) to the aluminum if the specific heat of aluminum is $0.215 \text{ cal/g } ^{\circ}\text{C}$.
- 7.2×10^5
 - 7.2×10^7
 - 7.2×10^3
 - 7.2×10^1
 - 7.2×10^2
13. An 8 000-kg aluminum flagpole 100-m high is heated by the sun from a temperature of 10°C to 20°C . Find the increase in internal energy (in J) of the aluminum. (The coefficient of linear expansion is $24 \times 10^{-6} (\text{^{\circ}\text{C}})^{-1}$, the density is $2.7 \times 10^3 \text{ kg/m}^3$, and the specific heat of aluminum is $0.215 \text{ cal/g } ^{\circ}\text{C}$.)
- 7.2×10^5
 - 7.2×10^7
 - 7.2×10^3
 - 7.2×10^1
 - 7.2×10^2
14. Two kilograms of water at 100°C is converted to steam at 1 atm. Find the work done (in J). (The density of steam at 100°C is 0.598 kg/m^3 .)
- 3.4×10^5
 - 1.2×10^5
 - 4.6×10^4
 - 2.1×10^4
 - 3.4×10^4

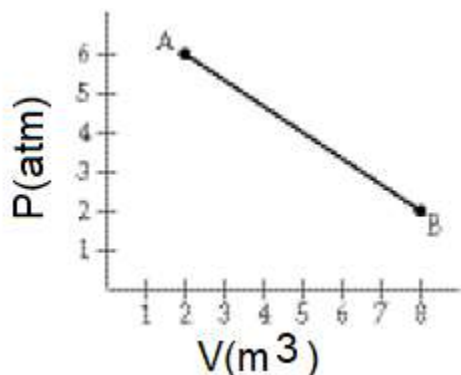
15. Two kilograms of water at 100°C is converted to steam at 1 atm. Find the change in internal energy (in J). ($L_V = 2.26 \times 10^6 \text{ J/kg.}$)
- a. 2.1×10^4
 - b. 4.5×10^6
 - c. 3.4×10^5
 - d. 4.2×10^6
 - e. 2.1×10^6
16. If an object feels cold to the touch, the only statement that you can make that must be correct is that
- a. the object has a smaller coefficient of thermal conductivity than your hand.
 - b. the volume of the object will increase while it is in contact with your hand.
 - c. the object contains less thermal energy than your hand.
 - d. the object is at a lower temperature than your hand.
 - e. the object cannot be a liquid.
17. Which of the following statements is correct?
- a. You only need to know the amount of thermal energy a body contains to calculate its temperature.
 - b. The temperature of a body is directly proportional to the amount of work the body has performed.
 - c. The quantity of thermal energy exchanged by two bodies in contact is inversely proportional to the difference in their temperatures.
 - d. The quantity of thermal energy exchanged by two bodies in contact is directly proportional to the difference in their temperatures.
 - e. Different amounts of thermal energy are transferred between two bodies in contact if different temperature scales are used to measure the temperature difference between the bodies.
18. In which process will the internal energy of the system **NOT** change?
- a. An adiabatic expansion of an ideal gas.
 - b. An isothermal compression of an ideal gas.
 - c. An isobaric expansion of an ideal gas.
 - d. The freezing of a quantity of liquid at its melting point.
 - e. The evaporation of a quantity of a liquid at its boiling point.

19. For an astronaut working outside a spaceship, the greatest loss of heat would occur by means of
- conduction.
 - convection.
 - radiation.
 - conduction and convection.
 - conduction and radiation.
20. Which statement below regarding the First Law of Thermodynamics is most correct?
- A system can do work externally only if its internal energy decreases.
 - The internal energy of a system that interacts with its environment must change.
 - No matter what other interactions take place, the internal energy must change if a system undergoes a heat transfer.
 - The only changes that can occur in the internal energy of a system are those produced by non-mechanical forces.
 - The internal energy of a system cannot change if the heat transferred to the system is equal to the work done by the system.
21. How much heat, in joules, is required to convert 1.00 kg of ice at 0°C into steam at 100°C ? ($L_{\text{ice}}=333\text{J/g}$, $L_{\text{steam}}=2.26\times 10^3\text{J/g}$)
- 3.35×10^5
 - 4.19×10^5
 - 2.36×10^6
 - 2.69×10^6
 - 3.01×10^6
22. Water at room temperature, 20°C , is pumped into a reactor core where it is converted to steam at 200°C . How much heat (in J) is transferred to each kilogram of water in this process? ($C_{\text{steam}}=2010\text{J/Kg }^{\circ}\text{C}$, $L_{\text{steam}}=2.26\times 10^3\text{J/g}$, $1\text{ cal} = 4.186\text{ J}$.)
- 3.35×10^5
 - 7.53×10^5
 - 2.67×10^6
 - 2.80×10^6
 - 3.01×10^6

23. A gas expands from A to B as shown in the graph. Calculate the work (in joules) done by the gas. (1 atm = $1.01 \times 10^5 \text{ N/m}^2$.)



- a. 12
 - b. 24
 - c. 1.21×10^6
 - d. 2.42×10^6
 - e. 3.64×10^6
24. A gas expands as shown in the graph. If the heat taken in during this process is $1.02 \times 10^6 \text{ J}$ and 1 atm = $1.01 \times 10^5 \text{ N/m}^2$, the change in internal energy of the gas (in J) is



- a. -2.42×10^6
 - b. -1.40×10^6
 - c. -1.02×10^6
 - d. 1.02×10^6
 - e. 1.40×10^6
25. In a thermodynamic process, the internal energy of a system in a container with adiabatic walls decreases by 800 J. Which statement is correct?
- a. The system lost 800 J by heat transfer to its surroundings.
 - b. The system gained 800 J by heat transfer from its surroundings.
 - c. The system performed 800 J of work on its surroundings.
 - d. The surroundings performed 800 J of work on the system.
 - e. The 800 J of work done by the system was equal to the 800 J of heat transferred to the system from its surroundings.

26. If a person in Alaska were locked out of his house on a day when the temperature outside was -40°C , he would be most likely to lose the most thermal energy by
- conduction.
 - convection.
 - radiation.
 - all of the above.
 - convection and radiation.
27. The Earth intercepts $1.27 \times 10^{17} \text{ W}$ of radiant energy from the Sun. Suppose the Earth, of volume $1.08 \times 10^{21} \text{ m}^3$, was composed of water. How long would it take for the Earth at 0°C to reach 100°C , if none of the energy was radiated or reflected back out into space?
- 26.9 y
 - 113 y
 - $2.69 \times 10^4 \text{ y}$
 - $1.13 \times 10^5 \text{ y}$
 - $2.69 \times 10^7 \text{ y}$
28. A team of people who traveled to the North Pole by dogsled lived on butter because they needed to consume 6,000 dietitian's calories each day. Because the ice there is lumpy and irregular, they had to help the dogs by pushing and lifting the load. Assume they had a 16-hour working day and that each person could lift a 500 N load. How many times would a person have to lift this weight 1.00 m upwards in a constant gravitational field where $g=9.8 \text{ m/s}^2$ to do the work equivalent to 6,000 Calories?
- 50.2
 - 492
 - 5130
 - 50,200
 - 492,000
29. Duff states that equal masses of all substances have equal changes in internal energy when they have equal changes in temperature. Javan states that the change in internal energy is equal to a constant times the change in temperature for every ΔT , no matter how large or how small ΔT is, but that the constant is different for different substances. Which one, if either, is correct?
- Neither, because the specific heat depends on the substance and may vary with temperature.
 - Neither, because a change of state may involve release or absorption of latent heat.
 - Neither because a substance may do work during a temperature change.
 - All of the statements above are correct.
 - Only statements (a) and (b) are correct.

- 30.** Which of the following statements is(are) correct when an ideal gas goes from an initial to a final state in a single process?
- a.** No work is done on or by the gas when the volume remains constant.
 - b.** No energy is transferred into or out of the gas as heat transfer when the temperature remains constant.
 - c.** The internal energy of the gas does not change when the pressure remains constant.
 - d.** All the statements above are correct.
 - e.** Only statements (a) and (b) above are correct.
- 31.** Beryl states that insulation with the smallest possible thermal conductivity is best to keep a house warm in winter, but worst for keeping a house cool in summer. Sapphire insists the reverse is true: low thermal conductivity is good in the summer, but bad in the winter. Which one, if either is correct?
- a.** Beryl, because low thermal conductivity results in low heat transfer.
 - b.** Beryl, because low thermal conductivity results in high heat transfer.
 - c.** Sapphire, because low thermal conductivity results in low heat transfer.
 - d.** Sapphire, because low thermal conductivity results in high heat transfer.
 - e.** Neither, because low heat transfer is desirable both in summer and in winter.