Chapter 1

Fundamentals of Refrigeration

Refrigeration, regardless of the means by which it is obtained, may be defined as *a process of the removal of heat from a substance or from a space*. Thus, if heat is removed from any substance or space, that substance or space becomes cool. In making the statements *warm* or *cool,* you must remember that these are relative statements. What is warm or cool is relative to surrounding temperatures. Thus, you must consider temperatures as measured by means of thermometers to arrive at the relationship between heat and cold.

From the foregoing statement, it follows that the problem of refrigeration concerns itself with heat removal from a substance or a space. Heat is a form of *energy,* such as electrical, chemical, and mechanical energy. It must be stated that energy can be neither created nor destroyed, but can be converted from one form to another, with temperature indicating its relative intensity.

The first method of refrigeration in recent history was by means of cutting and storing ice in cabinets. Later, considerable progress was made in the commercial manufacture of artificial ice. Probably the most widely used form of refrigeration was the manufacture of ice by means of an electrically driven compressor (with ammonia used as the refrigerant), from which 100-lb cakes of ice were produced commercially. The development of household refrigeration followed, and since 1923, this type of refrigeration machine has been manufactured in large quantities and has become commercially successful. Today, this type of machine is no longer considered a luxury; it is considered a necessity by all of us. From this effort, we further expanded into the commercial refrigeration and the frozen foods fields.

Applications of Refrigeration

Food preservation is the largest application of mechanical refrigeration today. Without modern refrigerating machinery, the packinghouse industry, transportation of perishable foods, and preparation of many other edibles would not only be difficult, but, in many cases, almost impossible. Refrigeration has not only saved quantities of meat, fish, eggs, milk, and cream from spoilage, it has also played an important part in the diet revision of the world. No longer are the inhabitants of one hemisphere, country, or locality dependent upon local foods; they may draw upon the entire world as a resource. This has made possible great developments in agriculture and livestock raising in countries far distant from potential markets. It has permitted the utilization of the great productivity of the tropics in supplying fruits and foods for other parts of the world.

Nearly all of the industries involved in the preparation of foods and drinks make extensive use of artificial ice or mechanical refrigeration. The dairy industry, for example, finds that for precooling milk and cream and in the manufacture of butter and ice cream, refrigeration is indispensable. In many industries, cooling and conditioning of air is an important phase of the manufacturing process. Among these processes are the manufacture of photographic films, explosives, machine production of cigars and cigarettes, candy and chewing gum, and rayon. They are important users of refrigeration, not only because of their size, but also because of their dependence on conditioned air for successful operation.

The cooling of liquids plays an important part in the mechanical and chemical industries. Oil-tempering baths are kept at constant low temperatures through refrigeration. In the jackets of nitrators and mixing machines for the celluloid, smokeless powder, and rubber industries, cooled brines are used to keep the mixtures at correct temperatures. In the chemical industry, use of refrigeration for drying air and gases and for purifying solutions by means of crystallization at definite temperatures is steadily increasing.

Physical Units

To obtain a clear conception of the functioning of a mechanical refrigeration system, it is imperative to consider and understand the physical and thermal properties underlying the production of the artificial cold.

Practical Dimensions

Although most of us in this country use the English system of measurement, we must also consider the metric system in our discussions. You will notice that many items purchased in stores give weights and measures in both the English and the metric systems. A comparison will now be made between the two systems, but this book will use primarily the English system (Figure 1-1). From the following information, you may readily convert from one system to another.

Table 1-1 shows prefixes commonly used in the metric system. Table 1-2 shows metric equivalents.

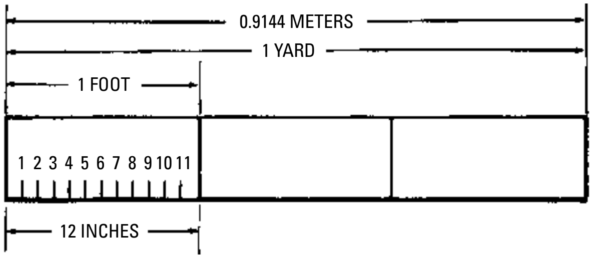


Figure 1-1 Relationship of centimeters to feet and meters to yards.

Table 1-1 Metric Prefixes

Prefix	***Meaning***
Micro	one millionth
Milli	one thousandth
Centi	one hundredth
Deci	one tenth
Deca	ten
Hecto	one hundred
Kilo	one thousand
Mega	one million

Table 1-2 Metric Equivalents

Measurement	***Equivalent***
10 millimeters	1 centimeter
100 centimeters	1 meter
1000 millimeters	1 meter
25.4 millimeters	1 inch
2.54 centimeters	1 inch
30.48 centimeters	1 foot
304.8 millimeters	1 foot
3.28083 feet	1 meter
39.36996 inches	1 meter
0.62137 miles	1 kilometer
1.60935 kilometers	1 mile
0.9144 meters	1 yard

Square Measure

Table 1-3 shows common units of area, covering both the English and the metric systems (Figure 1-2).

Table 1-3 Common Units of Area

Unit	*Equivalent*
1 square foot	144 square inches
1 square yard	9 square feet
1 square foot	929.03 square centimeters
1 square yard	0.836 square meters
1 foot	30.48 centimeters
1 yard	0.9144 meters

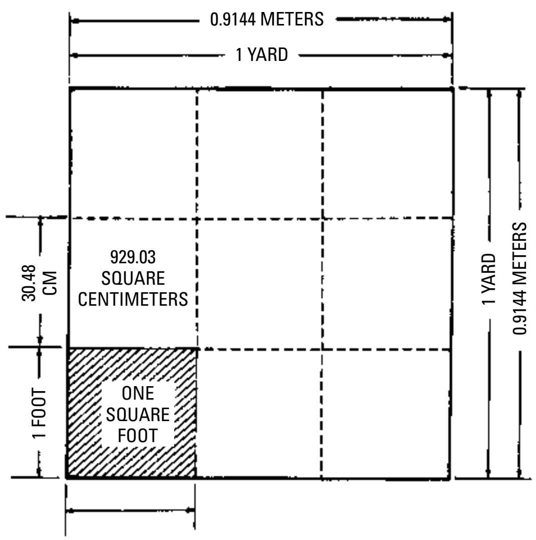


Figure 1-2 Relationship between various metric and English square units.

Cubic Measure or Measure of Solids

Table 1-4 shows common measurements of three-dimensional spaces of volume in both the English and the metric systems. (Figure 1-3).

Table 1-4 Common Measurements of Three-Dimensional Space

Unit	*Equivalent*
1 cubic foot	1728 cubic inches
1 cubic yard	27 cubic feet
1 cubic foot	28,316.84659 cubic centimeters
1 cubic foot	0.02832 cubic meters
1 cubic yard	0.7646 cubic meters

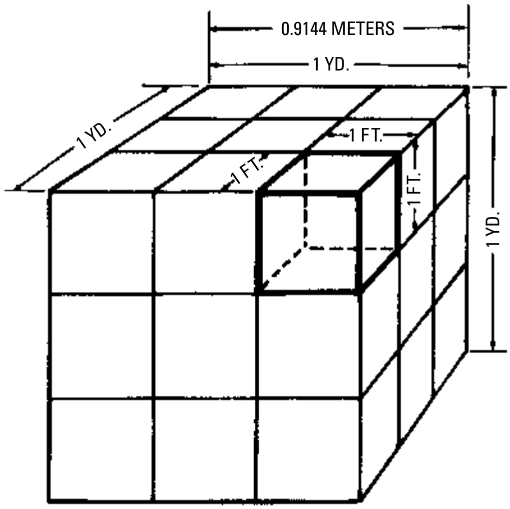


Figure 1-3 Relative amount of space occupied by a cubic Foot as compared to a cubic yard.

Mass

The quantity of matter that a body contains is called its *mass*. The space a body occupies is called its *volume*. The relative quantity of matter contained in a given volume is called its *density*. The relation of volume to mass is expressed by the term density and is written as follows:

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

$$\text{Mass} = \text{density} \times \text{volume}$$

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

The relative density of a substance is the ratio of its density to the density of pure water at a temperature of approximately 39.2°F (Fahrenheit) or 4°C (Celsius, formerly centigrade).

Mass is a property of all matter. Gas, air, water, and metals have mass. The weight of a substance is caused by the Earth's attraction on a substance (gravity). The only condition in which a substance has no weight is when it is falling in a perfect vacuum under the influence of gravity. At all other times it has weight. Therefore, weight divided by acceleration due to gravity is mass:

$$g = \text{ acceleration due to gravity} = 32.2 \text{ ft/s}^2$$

The unit of mass is the *slug*. The definition of a slug is 32.74 pounds, or 14.505 kilograms. Our units of weight are the ounce, the pound, and the ton:

$$16 \text{ ounces} = 1 \text{ pound}$$

$$2000 \text{ pounds} = 1 \text{ ton}$$

For example, to determine the mass of a 10-lb substance, use the following equation:

$$\text{Mass} = \frac{W}{g}$$

where the following is true:

W = weight of substance

g = acceleration due to gravity (32.2 ft/s^2)

$$\text{Mass} = \frac{10}{32.2} = 3.22 \text{ slugs}$$

Table 1-5 shows the metric units of weight and their conversion to the English system.

Before defining specific gravity, it should be explained that there must be a certain temperature used in arriving at specific gravity because liquids, solids, and gases will expand when heated and contract when cooled. In the preceding discussion on density of water, 39.2°F (4°C) was used.

Specific Gravity

Specific gravity is the ratio between the weight of a given volume of any substance and the weight of the same volume of some other substance taken as a standard. For solids and liquids, this standard is distilled water at a temperature of 39.2°F. Figure 1-4 gives the

Table 1-5 Metric Units and English Equivalents

Metric Unit	*English Equivalent*
1 grain	0.6480 grams
1 ounce	28.3495 grams
1 pound	0.45359 kilograms
1 short ton	0.9718 metric tons
1 long ton	1.01605 metric tons
1 gram	15.4324 grains
1 gram	0.03527 ounces
1 kilogram	2.20462 pounds
1 metric ton	1.10231 short tons
1 metric ton	0.98421 long tons

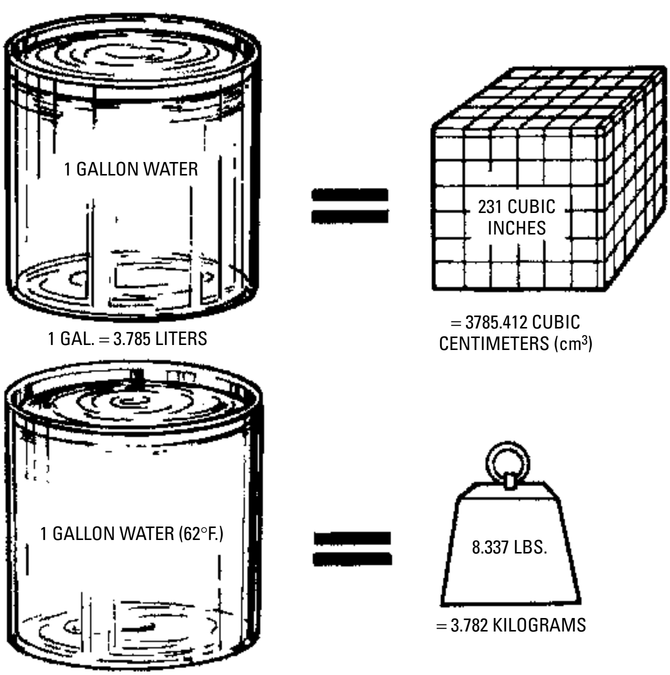


Figure 1-4 Difference between volume and weight as shown by a 1-gallon unit of water at 62°F.

volume and weight of 1 gallon of water at 62°F. In Figure 1-5, the weight of 1 in^3 of water is compared with 1 ft^3.

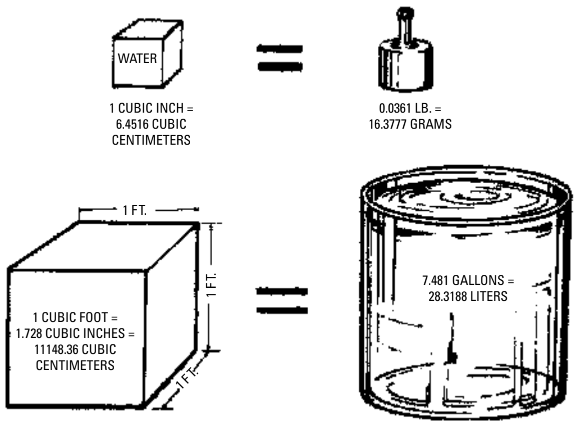


Figure 1-5 Comparison of the weight of 1 in^3 of water and the volume of 1 ft^3 and their metric equivalents.

For gases and vapors, some substance must be taken that is itself a gas, and so the standard for air is *hydrogen*. For solids and liquids, therefore,

$$\text{Specific gravity} = \frac{\text{weight of body in air}}{\text{weight of same volume of water}}$$

$$= \frac{\text{weight of body in air}}{\text{loss of weight in water}}$$

Figure 1-6 shows how to determine the specific gravity of a piece of cast iron. It will be noted that its weight in air equals 15 lb, and the weight registered by the same scale when the pieces are totally submerged in water is 12.9 lb. Substituting in the formula,

$$\text{Specific gravity} = \frac{W}{W - w} = \frac{15}{15 - 12.9} = 7.14$$

Atmospheric, Absolute, and Gage Pressures

It is of the utmost importance that refrigeration students understand the meaning of the various kinds of pressure as related to refrigeration.

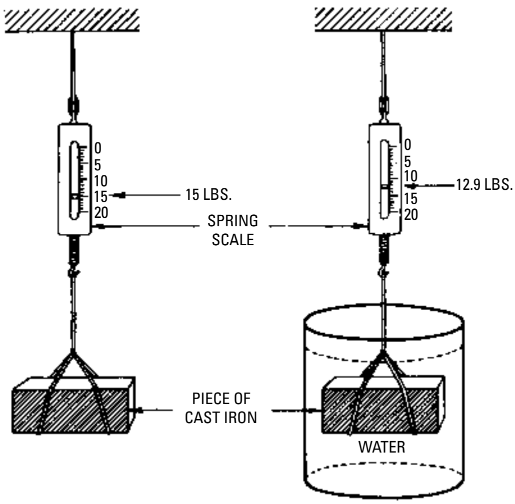


Figure 1-6 Comparison of the specific gravity of a piece of cast iron in air and in water.

Atmospheric Pressure

Atmospheric pressure is pressure that is exerted by the atmosphere in all directions, as indicated by a barometer. Standard atmospheric pressure is considered to be 14.695 pounds per square inch (usually written 14.7 psi), which is equivalent to 29.92 inches of mercury (in Hg).

When using 14.7 psi or 29.92 in Hg as atmospheric pressure, bear in mind that this is the perfect atmospheric pressure at sea level and will vary (plus or minus) with altitude and local atmospheric pressures. The temperature used for determining standard atmospheric pressure is 62°F.

As the altitude increases above sea level, the barometric pressure decreases. This is mentioned because in working with refrigeration, you will find later that vacuum measurements must be made. Altitude also affects the boiling point of the refrigerants. To show the results, pressures are given in Table 1-6 for various altitudes.

Barometers should always be corrected to the sea-level reading, but from the preceding figures, the actual barometric pressure can be found for the altitude at which you are working. This is a standard reading and does not take into account high and low atmospheric pressures that occur with weather changes.

Table 1-6 Effects of Altitude on Atmospheric Pressure

Altitude in Feet	*Psi*	*Inches of Mercury*
Sea Level = 0	14.72	29.92
1000	14.17	28.80
2000	13.64	27.73
3000	13.13	26.69
4000	12.64	25.69
5000	12.17	24.74
6000	11.71	23.80
7000	11.27	22.91
8000	10.85	22.05
9000	10.45	21.24
10,000	10.06	20.45
11,000	9.69	19.70
12,000	9.33	18.96
13,000	8.98	18.25
14,000	8.64	17.56
15,000	8.32	16.91

Absolute Pressure

Absolute pressure is the sum of gage pressure and atmospheric pressure at any particular time. For example, if the pressure gage at one particular time reads 53.7 lb, the absolute pressure would be 53.7 + 14.7, or 68.4 psi.

The aforementioned definitions may be written as follows:

$$\text{Absolute pressure} = \text{gage pressure} + 14.7$$

where 14.7 is the normal atmospheric pressure. Then

$$\text{Gage pressure} = \text{absolute pressure} - 14.7$$

Boyle's Law for Perfect Gases

Boyle's law refers to the relationship between the pressure and volume of a gas and may be stated as follows: *With temperatures constant, the volume of a given weight of gas varies inversely as its absolute pressure*. This is illustrated in Figure 1-7.

Boyle's law is written mathematically as

$$P_1 V_1 = P_2 V_2$$

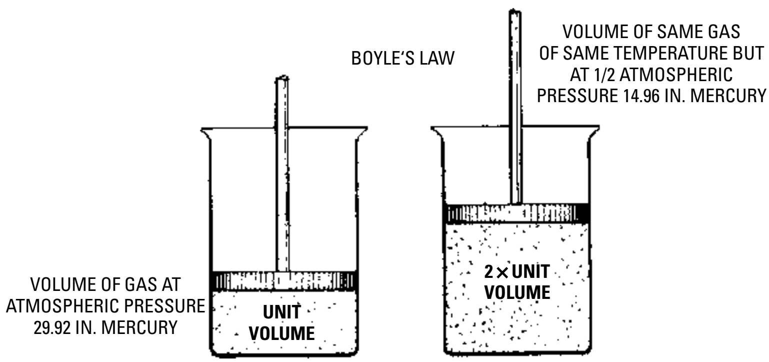


Figure 1-7 Boyle's Law for a perfect gas.

or

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

where the following is true:

P_1 = absolute pressure of a quantity of perfect gas before a pressure change
P_2 = absolute pressure after pressure change
V_1 = volume of gas at pressure P_1
V_2 = volume of gas at pressure P_2

Since $P_1 V_1$ for any given case is a definite constant quantity, it follows that the product of the absolute pressure and volume of a gas is constant:

$$PV = C \text{ (when temperature is kept constant)}$$

At this time it should be mentioned that any change in the pressure and volume of a gas at constant temperature is called an *isothermal change*.

Charles's Law

Charles's law refers to the relationship between the pressure, volume, and temperature of a gas and may be stated as follows: *At a constant pressure, the volume of a gas varies directly as the absolute temperature; at a constant volume, the pressure varies directly as the absolute temperature*. This is illustrated in Figure 1-8.

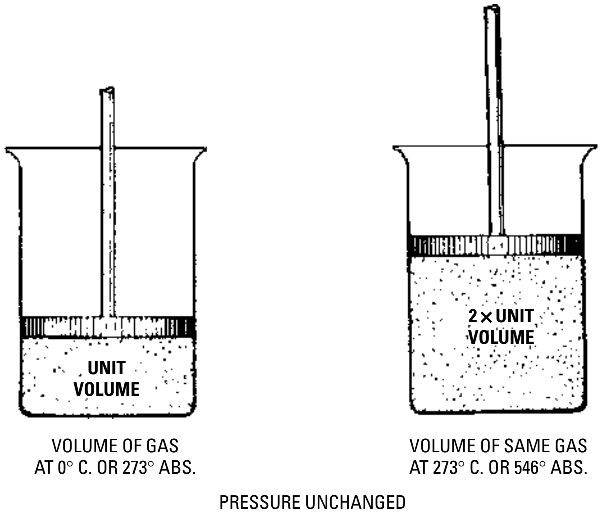


Figure 1-8 Charles's Law for a perfect gas.

When heat is added to a constant volume, the relation is written as follows:

$$P_1 T_2 = P_2 T_1$$

or

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

For the same temperature range at a constant pressure:

$$V_1 T_2 = V_2 T_1$$

or

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

Combined, these laws read as follows:

$$P_1 V_1 T_2 = P_2 V_2 T_1$$

or

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

Since volume is proportional to weight, the relationship between P, V, and T for any given weight of gas, W is

$$PV = WRT$$

where the following is true:

P = absolute pressure of gas (in points)
V = volume of gas
W = weight of gas
R = constant, depending on gas under consideration
T = absolute temperature (°F)

Figure 1-9 shows the four standard temperature scales. A comparison of the absolute Celsius and Celsius scales shows the absolute scale has 0° at 273° below the freezing point of water, which is 0° on the Celsius scale. Thus, 0°C would be 273°C absolute, and 10°C would be 283°C absolute.

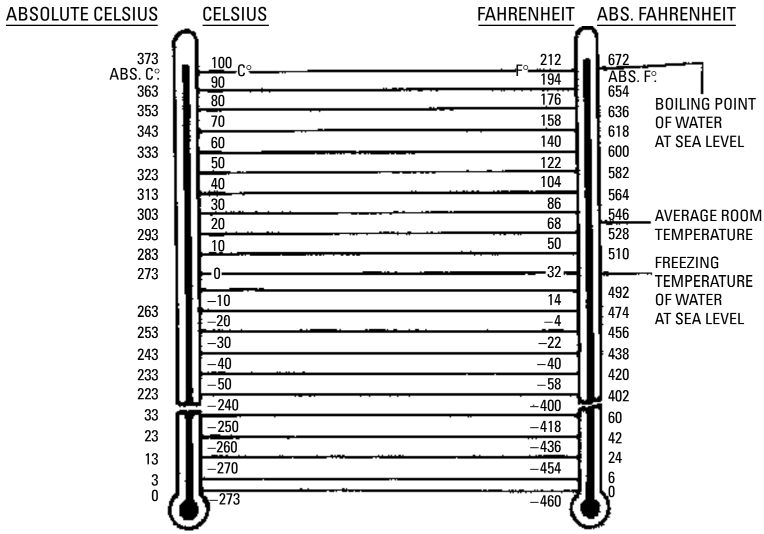


Figure 1-9 Four standard temperature scales.

Figure 1-9 also shows the Fahrenheit and absolute Fahrenheit scales. Freezing of water on the Fahrenheit scale would be 32°; on the absolute scale it would be 492°.

The methods and scales used to measure temperature have been chosen by scientists, and the following standards have been established. The common American scales are the Fahrenheit scale and the Fahrenheit absolute scale (Rankine scale), whereas the metric system embodies the Celsius and Celsius absolute scales (Kelvin scale).

The Fahrenheit scale is so fixed that it divides the heat level from the melting temperature of ice to the boiling point of water into 180 equal divisions and sets the melting point of ice at 32 divisions above the zero scale. Therefore, ice melts at 32°F, and water boils at 212°F (180° + 32°F), assuming the standard atmospheric pressure of 14.7 psi at sea level. The Fahrenheit absolute (F_a) scale uses the same divisions as the Fahrenheit scale, but sets zero at the temperature where molecular action of all substances ceases (absolute zero), where no heat exists in the body and the temperature cannot be lowered any further. This temperature corresponds to –460°F, and water boils at 672°F, assuming standard atmospheric pressure.

The Celsius scale has coarser divisions than the Fahrenheit scale, and the zero (0°) of this scale is set at the melting temperature of ice. The boiling point of water is fixed 100 divisions above that point, or 100°C, assuming atmospheric pressure to be 14.7 psi at sea level.

Occasionally you will have to convert Celsius reading to Fahrenheit reading and vice versa. To convert Celsius degrees into Fahrenheit degrees, use the following equation:

$$°F = \frac{9}{5}(°C + 32)$$

or

$$°F = 1.8\,(°C + 32)$$

To convert Fahrenheit degrees into Celsius degrees, use the following equation:

$$°C = \frac{5}{9}(°F - 32)$$

or

$$°C = 0.55555555\,(°F - 32)$$

To convert Fahrenheit degrees to Fahrenheit absolute, use the following equation:

$$°F_a = (°F + 460)$$

To convert Celsius degrees into Celsius absolute, use the following equation:

$$°C_a = (°C + 273)$$

Power and Work

Power may be defined as the capability of performing mechanical work as measured by the rate at which it is or can be done. Stated in simple language, power is *the rate of doing work*. If a horse pulls a weight of 220 lb against gravity at the rate of 2.5 ft/s (as shown in Figure 1-10), and if the friction in the pulling arrangement is neglected, the horse will develop one horsepower (1 hp). The work done in a given time divided by the units of time gives *the average rate of doing work,* or the *power,* which may be written as follows:

$$\text{Power} = \frac{\text{work}}{\text{time}}$$

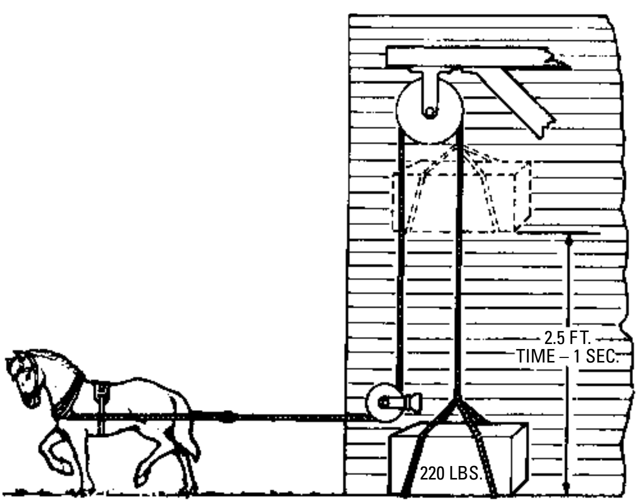


Figure 1-10 One horsepower, using a horse to lift.

Units of Power

The most commonly used units of power are the foot-pounds per unit time and horsepower. The horsepower is the rate of doing work at 550 foot-pounds per second (ft-lb/s), or 33,000 foot-pounds per minute (ft-lb/min). It is equivalent to the power used to raise a weight of 33,000 pounds against gravity at a rate of one foot per minute:

$$\text{hp} = \frac{\text{foot-pounds}}{33{,}000 \times \text{time (minutes)}}$$

Figure 1-11 shows another way of illustrating one horsepower. The motor is said to develop 1 hp when the 550-lb weight is lifted against gravity at a rate of 1 ft/s. If a weight of 3000 lb is lifted through a distance of 35 feet in 2 minutes, the required number of horsepower would be the following:

$$\text{hp} = \frac{3000 \times 35}{33{,}000 \times 2} = 1.59$$

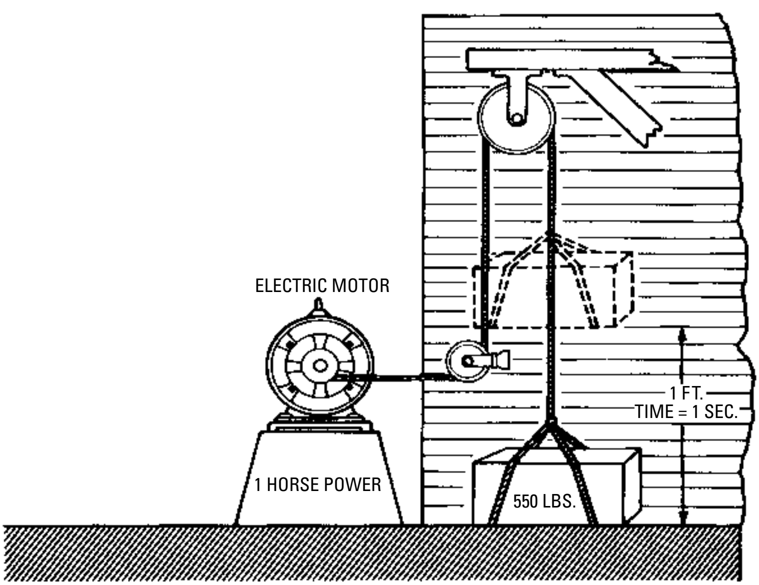


Figure 1-11 One horsepower, using a motor to lift.

Energy

A body is said to possess energy when it can do work. Energy exists in various forms, such as the following:

- Mechanical energy
- Electrical energy
- Heat energy (thermal)

Other well-known forms of energy are *kinetic* and *potential energy*.

Energy may be transmitted from one form to another by various processes (such as mechanical, thermal, chemical, or electrical), but the energy lost or gained is either kinetic or potential. If the process is mechanical (such as impact, compression, or the application of a mechanical force), work is done. If it is thermal, radiation, or conduction, then heat is added or withdrawn. Very simple relations exist between mechanical, electrical, and heat energy, and they may be written as shown in Table 1-7.

Table 1-7 Energy Equivalents

Unit	*Equivalent*
778 foot-pounds (ft-lb)	1 British thermal unit (Btu)
2546 British thermal units (Btu)	1 horsepower-hour (hp-hr)
39,685 British thermal units (Btu)	1 kilogram-calorie (kg-cal)
746 watts (W)	1 horsepower (hp)

Other equivalents to 1 Btu are as follows:

0.001036 lb water evaporated at 212°F

0.0000688 lb carbon oxidized

0.000393 hp-hr

107.6 kg-m

778 ft-lb

1055 W-s

From these equivalents, it may immediately be apparent that heat is an entity, a real something, for it has a unit of measure and each unit is convertible to other forms of energy at a constant rate of exchange. Since refrigeration is an art that is concerned with the heat problem, those who desire a thorough understanding of the refrigeration process must know the simple facts concerning heat.

Theory of Heat

The technical nature of heat was not understood for a long time, and several different theories were advanced to explain it. The concept that is generally accepted is the so-called *molecular theory* because it is based on the theory that all matter is composed of innumerable, separate, and minute particles called *molecules*. The molecule is composed of two or more atoms and is so small that an extremely powerful microscope is required to see it.

Molecular Theory

The molecular theory is based on the supposition that the molecules of a substance are not attached to each other by any bond or cement, but are held together by a force known as *cohesion,* or *mutual attraction,* a phenomenon somewhat similar to the attraction offered by a magnet for steel particles. The molecules, however, are not physically bound or in contact with each other (like the iron filings) but are actually separated to such an extent in some instances (and usually so at more elevated temperatures) that the space separating two adjoining molecules is larger than either particle. Furthermore, the molecules are not fixed or stationary but revolve and vibrate within the orbits or limits of their allotted space.

Each substance on Earth is composed of different ingredients or various combinations of molecules, and, therefore, each particular kind and mixture, together with additional peculiarities in physical assembly, has a structure differing from other materials. It is the rapidity of the motion of the molecules of a body or mass that determines the intensity of its heat. As a specific example of the connection between the theory of heat and the molecular theory, a bar of lead may be used for illustration. When cold, the lead molecules undergo comparatively slow motion. But if the bar is heated, the molecular activity becomes more rapid. It continues until its temperature has been increased to a point where molecular motion becomes so rapid and the individual particles are so far separated from each other that, with further assimilation of heat, they become so weakly cohesive they can no longer hold the body in a rigid or solid mass and it devolves into a liquid.

Further application of heat forces the molecules to greater separation and speeds up their motion to such an extent that the liquid becomes more mobile and volatilization finally results, so that a gas or vapor is produced. The vapor thus formed no longer has a definite volume, such as it had in either the solid or molten (liquid) form, but will expand and completely fill any space that is provided for it. The vapor, of course, contains the same number of molecules that were in the original solid, but molecular action is extremely active in the vapor phase, comparatively slow in the liquid, and slowest in the solid form. The motion of the molecule is somewhat like a bicycle rider: at a slow speed, a very small orbit can be adhered to, but with each advance in the rate of travel, larger orbits are required.

To one learning the theory of heat and matter for the first time, it may appear extremely complex. Whether the molecular theory herein described (which is the one accepted by leading scientists) or any other theory is used as a basis of explanation, the facts regarding heat as we know it are in agreement with the molecular theory and from all tangible proof substantiate the theory in all respects. One of the simplest indications of value in support of the molecular theory is the age-old knowledge concerning expansion and contraction, caused by heat and cold, respectively, or, as we look at it from the molecular theory, the rate of molecular activity.

Temperature

Heat and temperature are closely allied, but it must be remembered that heat is energy itself representing the kinetic activity of the molecules composing a substance, whereas temperature is but a measure of the condition of a mass or body as it affects our ideas of warmth or cold. Heat is convertible into electrical, chemical, or mechanical energy.

Our ideas of cold and hot are merely relative ranges of temperature that affect our sense of feeling, which we term hot and cold. In a way, heat itself may be likened to water, for, of its own accord, it will flow only downhill; that is, it will pass from the hotter (or higher) range of temperature to the colder (or lower) plane.

Two miscible liquids poured together will find their own temperature level, as will a solid and a liquid, or even two solids in intimate contact. For example, if an ice cube is dropped into a tumbler of tepid water, it will be found that the colder substance will be able to assimilate a certain quantity of heat from the warmer substance. Thus, through this extraction, we find that the water has been cooled or reduced in temperature. Two substances of different temperatures in intimate contact tend to reach an equilibrium or balance by the dispersion of heat by one body and assimilation of heat by the colder body. If a thermometer is placed in contact with the substance, the degree or level of its temperature (or the measure of the intensity of its heat) will be indicated.

Thermometers

The instrument in common use for measuring temperature, known as the *thermometer,* operates on the principle of the expansion and contraction of liquids (and solids) under varying intensities of heat. The ordinary mercury thermometer operates with a fair degree of accuracy over a wide range. It becomes useless, however, where temperatures below −38°F are to be indicated because mercury freezes at that point. Some other liquid, such as alcohol (usually colored for easy observation), must be substituted. The upper range for mercurial thermometers is quite high (about 900°F), and so it is apparent that for ordinary service and general use the mercury thermometer is usually applicable.

The operation of the thermometer depends on the effect of heat on the main body of mercury or alcohol contained in a bulb or reservoir. The liquid will expand or contract (rise or drop) in the capillary tube, which is inscribed with the various increments of an arbitrary scale, as shown in Figure 1-12. Several temperature scales are in existence and are used in various countries. The English, or Fahrenheit, scale is commonly used in the United States. Since the Celsius scale is so widely used in scientific work in all countries, an illustration of the comparison of thermometers is presented so that any one scale may be converted to another.

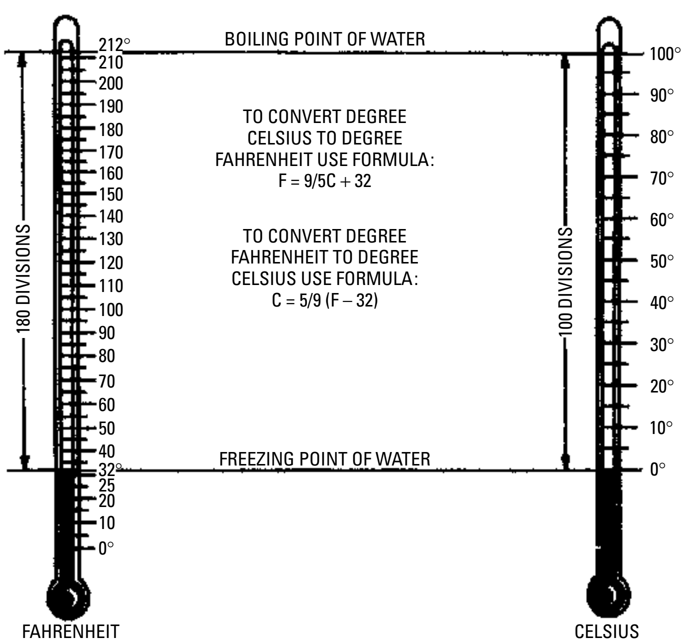


Figure 1-12 Relationship between fahrenheit and celsius scales used on thermometers.

Freezing on the Fahrenheit scale is fixed at 32°; on the Celsius scale, freezing is placed at 0°. On the Fahrenheit scale, the boiling point of pure water under the normal atmospheric pressure encountered at sea level is 212°; on the Celsius scale, under the same conditions, the boiling point is 100°.

In Figure 1-9, all four temperatures are shown; only the Fahrenheit and Celsius scales are shown in Figure 1-12 because these two scales are used most often.

As previously stated, temperature is a measurement of the intensity of the heat contained in a body, accomplished by means of a thermometer, just as a yardstick is used to measure the length of a body. However, a thermometer will not indicate the quantity or amount of heat contained in a body.

Absolute Zero

In the study of thermometer scales, the question of thermal limits is naturally considered. From the molecular theory, we are led to believe that with the removal of heat, molecular action is slowed down accordingly. It must naturally follow that, at some point, all heat will be removed and molecular activity will cease entirely.

The calculation in Figure 1-13 indicates that absolute zero is attained at a temperature of 460° below zero on the Fahrenheit scale. Temperatures as low as this have never been reached, although in some instances it has been approached within a few tenths of one degree. Bodies subjected to extremely low temperatures take on characteristics entirely different from those exhibited under normal conditions.

This calculation in Figure 1-13 is based upon the fact that atmospheric air expands 1.366 of its volume on being heated from 32° to 212°F. Therefore, it can be assumed inversely that if all possible heat were withdrawn, the volume of air would shrink to zero. With reference to our diagram, the volume of air at 32°F = 1. It follows that the volume of air at 212°F = 1.366. Constructing our triangles as shown, the following relations are obtained:

$$\frac{32 - X}{212 - X} = \frac{1}{1.366}$$

$$X = \frac{1.366 \times 32 - 212}{0.366}$$

$$= 460°\text{F}$$

$$(-459.62°\text{F Rounded to } -460° \text{ here})$$

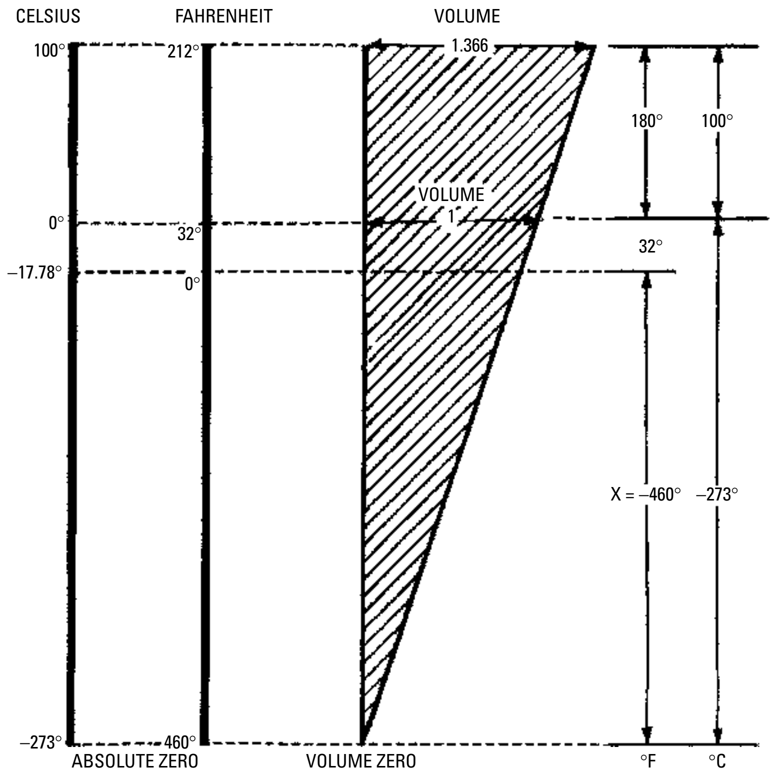
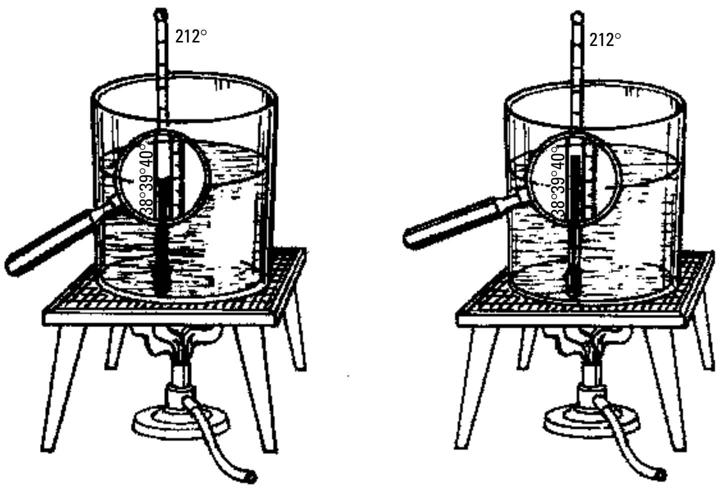


Figure 1-13 How absolute-zero temperature may be determined.

Units of Heat

The quantity of heat contained in a body is not measurable by a thermometer, which indicates only the temperature or intensity. For example, a gallon of water and a pint of water may have the same temperature, but we are certainly aware that the larger body must contain more heat as energy than the smaller body.

The unit of heat measure employed in this country is called the *British thermal unit* (more commonly referred to as the Btu). The Btu is that quantity of heat required to be added to one pound of pure water initially at a temperature of its greatest density (39°F) to raise its temperature one degree on the Fahrenheit thermometer (in this case, to 40°F). Roughly, a Btu may be said to be the quantity of heat required to raise the temperature of one pound of water one degree on the Fahrenheit scale, as shown in Figure 1-14.



(A) 1 lb water at 39°. (B) 1 lb water at 40°.

Figure 1-14 British thermal unit (Btu).

Just as the thermometer is used as a measure of intensity of heat of a body, the Btu is used to represent the quantitative energy. For example, one body at 50°F may contain twice as many Btu as another body at the same temperature because the bodies may be different in size or weight and, of great importance, may have different capacities for absorbing heat.

Specific Heat

It could be said that every substance on the Earth has a different capacity for absorbing heat. Some identical materials (especially natural formations) even give different values for samples secured in different localities. As an illustration of the difference in the heat capacities of materials, consider a pound of iron and a pound of water, both at 80°F. If the heat is removed from these bodies and the number of Btu extracted in cooling each mass to the same temperature is recorded, it will be found that each will give up a different amount. If each body had been at a temperature of 80°F and was cooled to 60° (a reduction of 20° on the Fahrenheit scale), it would be found that 1 lb of water at 80° cooled to 60° would evolve 20 heat units or 20 Btu.

If the iron (of the same weight) is cooled over the same range, only 2.6 Btu will be extracted (approximately $^1/_8$ of that taken from the same weight of water under identical conditions). It can be concluded, then, that all materials absorb heat in different capacities. By comparing the heat-absorbing qualities with a standard, we have a standard of measure, a gage to compare all substances. This measure is the amount of heat (expressed in Btu) required by one pound of a substance to change its temperature one degree Fahrenheit. Since water has a very large heat capacity, it has been taken as a standard; and since one pound of water requires one Btu to raise its temperature one degree, its rating on the specific-heat scale is 1.00. Iron has a lower specific heat, its average rating being 0.130; ice is 0.504; air is 0.238; and wood is 0.330. The more water an object contains, as in the case of fresh food or air, the higher the specific heat. Materials usually stored in a refrigerator have a high specific heat, averaging about 0.80.

Table 1-8 lists the average specific heats of common substances, comparing their relative heat capacities with water. Observe that metals have limited heat-storing powers as compared with water. This is one of the reasons why scalds from hot water burn so deeply; the water contains so much heat energy that a considerable amount is released and causes a worse burn than molten metal at a much higher temperature.

Table 1-8 Average Specific Heats

Material	*Specific Heat*	*Material*	*Specific Heat*
Water	1.000	Pine	0.650
Copper	0.900	Strong brine	0.700
Vinegar	0.920	Oak	0.570
Alcohol	0.659	Ice	0.504
Air	0.238	Glass	0.194
Mercury	0.333	Iron	0.130
Coal	0.241	Sulfur	0.202
Brass	0.094	Zinc	0.095

In Table 1-8 various substances are listed with their average specific values. By finding this factor, the amount of heat in Btu to be added or taken from a substance of known weight to bring about a change of one degree in its temperature may be calculated. Specific heat problems may be calculated easily by use of the following formula:

$$\text{Btu} = \text{Sp. heat} \times W(t_2 - t_1)$$

where the following is true:

$$W = \text{weight of the substance (in lb)}$$
$$\text{Sp. heat} = \text{specific heat of the substance to be heated}$$
$$t_2 - t_1 = \text{temperature change (°F)}$$

Where very accurate scientific work is done, certain allowances are made for the fact that the specific heat of a substance does not remain constant throughout the entire temperature range. Because the difference is not appreciable (except over a wide range), the refrigeration engineer usually regards the specific heat factor as a constant. Other materials (such as liquids and gases) also have specific heats. However, the calculations concerning the heat capacities of gases are further complicated by the pressures as well as the varying temperatures imposed. Despite the tendency to ignore the variance in specific heat and employ a constant, the values of water are given in Table 1-9.

Table 1-9 Specific Heats of Water (Value at 55°F, Taken as Unity)

Temp., °F	*Specific Heat*	*Temp., °F*	*Specific Heat*
20	1.0168	140	0.9986
30	1.0098	160	1.0002
40	1.0045	180	1.0019
50	1.0012	200	1.0039
60	0.9990	220	1.0070
70	0.9977	240	1.0120
80	0.9970	260	1.0180
90	0.9967	280	1.0230
100	0.9967	300	1.0290

The specific heat of various foods and their containers are of interest to the refrigeration engineer in estimating the amount of heat to be extracted when cooling a refrigerator load. The heat that can be felt and detected is termed the *sensible heat*; its name denotes the heat we can sense or feel, preventing confusion with other heats. It is the sensible heat that forms the preponderance of the heat load the refrigerating machine is called on to remove in cooling most edibles. Of course, ice manufacturing plants and other storage warehouses that freeze edibles are additionally concerned with other heat problems.

Most foods have high water content and, therefore, a large heat capacity, averaging about 0.80. A few foods are listed in Table 1-10 so that the specific heats of various edibles can be ascertained and the truth of the statement regarding the water content verified. The specific heats listed are those of fresh foods before freezing.

Table 1-10 Specific Heats of Foods

Food	*Specific Heat*	*Food*	*Specific Heat*
Apples	0.92	Eggs	0.76
Beans	0.91	Fish	0.80
Beef	0.75	Grapes	0.92
Butter	0.60	Milk	0.90
Cabbage	0.93	Peaches	0.92
Cheese	0.64	Pork	0.50
Chicken	0.80	Potatoes	0.80
Celery	0.91	Veal	0.70
Cider	0.90	Watermelon	0.92

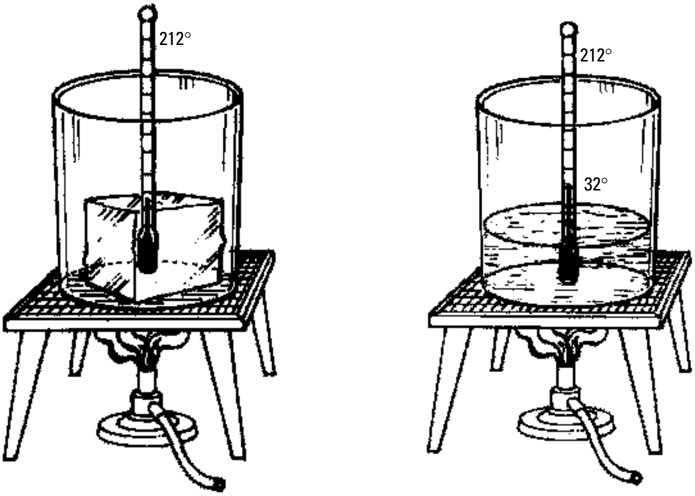
An estimate of the heat quantities that must be removed from the food and its containers can be easily accomplished. Take, for example, 1000 lbs of cider contained in glass bottles with a weight of 75 lb. The bottles in turn have been packed in pine boxes totaling 50 lb. The whole shipment arrived at a temperature of 80°F. The heat, in Btu, would have to be extracted to cool these materials to 50°F. The Btu can be calculated as follows:

1. 1000 lb of cider having a specific heat of 0.90, cooled over a 30° range $= 1000 \times 0.90 \times 30 = 27{,}000$ Btu.
2. 75 lb of glass with a specific heat of 0.194, cooled 30° $= 75 \times 0.194 \times 30 = 436.5$ Btu.
3. 50 lb of pine wood having a specific heat of 0.650 cooled 30° $= 50 \times 0.650 \times 30 = 975$ Btu, or a total of $27{,}000 + 436.5 + 975 = 28{,}411.5$ Btu.

By observing a thermometer immersed in the cider and taking an average reading, we could calculate at any time during the cooling process how much of the sensible heat had been removed. In fact, we could calculate any heat load of any material by finding the specific heat value and the cooling range. If we wanted to freeze the product (in this specific case, the cider), an additional factor would have to be considered, namely, the *latent heat*.

Latent Heat

One of the most mystifying laws to the layman is that of *latent heat*. The word *latent* expresses it aptly enough, for it means hidden or not apparent. To indicate clearly just what latent heat is, an example will be cited and an illustration will be shown in Figure 1-15.



(A) 1 lb ice at 0°F. (B) 1 lb water at 32°F.

Figure 1-15 Apparatus needed for conversion of 1 lb of ice at 0°F to water at 32°F.

We are aware that, under proper conditions, most substances are capable of assuming two or more physical states. For instance, lead, when cold, is a solid; when heated and molten, it is a liquid. Water is an outstanding example because it can assume three states—solid, liquid, and vapor—within a relatively short temperature range. Ice, of course, represents the solid state; water, the liquid; and steam, the vaporous, or gaseous, state.

Change of State

Temperature and heat play important parts in effecting changes from one state to another. For example, the only thing that would convert a tumbler of water to either of the other states (that is, a solid or a vapor) would be the addition or extraction of heat.

To prove the function that heat plays in effecting a change of state, let us take a block of ice and perform an experiment with it. To simplify the calculations, we will utilize a piece of ice that weighs exactly 1 lb with a thermometer frozen in the center of it, and we will assume that the block of ice is at a temperature of 0°F. We will not discuss the apparatus or computations necessary to measure the heat values applied but will assume the use of an imaginary Btu meter just as though such a device existed. Thus, every time a Btu of heat energy is expended, it will be registered on our meter.

Prior to beginning the experiment, we use a graph or a squared section of paper to plot the results. In the lower right-hand section of the paper (a part that will not be required for our log), we will put down the various findings. Setting up the apparatus is a simple matter. We require only a vessel and a source of heat, the latter being measurable by our meter. Place the piece of ice, weighing exactly 1 lb, in the vessel and indicate the start of the experiment by marking an *A* at the lower left-hand corner of the chart (as shown in Figure 1-16) to denote that the ice is at 0°F and that no heat has been added as yet. The specific heat of ice is not as great as that of water; in fact, it is just about half the value (0.504 to be exact). For this experiment let us assume it is 50 percent, or 0.5.

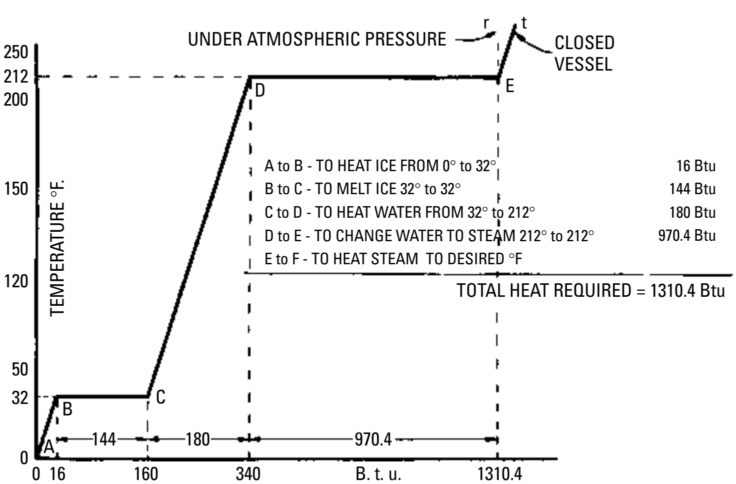


Figure 1-16 Heat units required to change 1 lb of ice at sea level and 0°F to steam at 212°F.

Now begin to add heat. According to its specific heat capacity, if $^1/_2$ Btu is added, the temperature of the ice increases 1°F. By adding heat and marking off the progress on the chart, it is found that the number of Btu required to bring the ice to the degree of sensible heat of 32°F would be

$$1\,(\text{lb of ice}) \times 0.5\,(\text{specific heat}) \times 32\,(\text{degrees rise}) = 16\,\text{Btu}$$

Thus, from *A* to *B* (a range of 32°) 16 Btu were required.

At 32°F, an interesting stage is reached. We find that the further addition of heat does not warm up the ice and the thermometer frozen in the test block of ice continues to indicate a sensible heat of 32°. Even providing the ice with a greater quantity of heat causes no observable change in temperature. The only noticeable feature is the melting of ice or conversion of a solid (ice) to a liquid (water). This important transformation is called the *change of state,* because of the fact that a solid is converted to a liquid. If we continued to add heat to the vessel containing the ice and water, we would find that when 144 Btu has been added, the entire block of ice would be converted to water and the water would have a temperature of 32°F. In other words, to change the state of 1 lb of ice at a temperature of 32°F to water at the same temperature requires 144 Btu. We will plot this on our chart as lines *B* and *C*.

Latent Heat of Fusion

We have had to apply 144 Btu, which was taken up by the ice at 32° and caused it to melt and assume its liquid state without a single indication of any rise or increase in its sensible heat. This masked assimilation of heat is termed *latent heat* and may be said to be the amount of heat units that must be supplied to a solid to change it to a liquid without an increase in temperature. It is readily understood that a reversal of the process, that is, for the liquid to assume its solid state (ice), heat must be extracted from water.

It requires the extraction of 144 Btu to cause 1 lb of water at a temperature of 32°F to freeze into a solid block of ice at 32°F. Every solid substance has a latent-heat value. The amount required to convert it or bring about a change of state is termed the *latent heat of fusion*. This heat (assimilated or extracted, as the case may be) is not measurable with the thermometer because the heat units are absorbed or expended in intermolecular work. Latent heat of fusion separates the molecules from their attractive forces so that a change of state is effected.

Latent Heat of Evaporation

Now let us refer again to the experiment we have under way. We have converted the solid to a liquid, both at 32°F, by adding 144 Btu. Having to deal with water, and knowing that it has a specific-heat factor of 1.00, we may look for a rise of one degree in temperature for each Btu added. This will hold until 212°F is indicated on the thermometer, assuming that the experiment is made at atmospheric pressure existing at sea level. Over the 180° range of 32 to 212°, we will be obliged to add 180 Btu to raise the temperature of 1 lb of water from the former to the latter temperature. This may be plotted as lines *C* and *D* on the chart, with the number of heat units recorded on the lower edge.

The further addition of heat would serve to bring out the fact that while the phenomenon of boiling would occur at 212°F, the temperature or sensible heat would not be increased. Just as ice requires a certain quantity of heat units (144 Btu/lb) to melt or convert it from a solid to a liquid without a rise in temperature, we find a similar condition existing. This time a liquid is being converted to a vapor, steam in this case.

Careful measurements have determined that the conversion of 1 lb of pure water at 212°F to steam at 212°F requires exactly 970.4 Btu when carried out at the normal pressure of the atmosphere encountered at sea level. If we carefully add heat and keep count of the Btu expended, we will find that when all the water has been changed to steam, 970.4 heat units will have been used. Thus, lines *D* and *E* may be plotted on the chart and a note of the number of Btu expended jotted down on the lower right-hand corner. The further addition of heat serves only to heat the steam, such as would be possible if it had been trapped or the experiment had been performed in a closed vessel so that heat could be applied to it.

Steam

Steam is the hot, invisible vapor state of water at its boiling point. The visible white vapor is really a collection of fine watery particles formed from true steam by condensation. Steam acts like all trite vapors or gases in that it has the following qualities:

- Fluidity
- Mobility
- Elasticity
- Exerts equal pressure in all directions

The difference in volume between water and steam at atmospheric pressure is 1646 to 1, and this wide ratio is manifested by nearly all gases and vapors. The heating of the steam generated in our experiment could be represented by lines *E* and *F* and would be extended over a considerable range. By referring to the log of our experiment, we can trace the heat quantities required to convert ice to water and steam and the action of heat on both liquid and solid. A tabulated form in the lower right-hand portion of the log represents the quantities of heat energy required to effect a change of state, rise in temperature, and other values.

From this experiment, it is apparent that heat added to a substance either increases its temperature or changes its state. The heat that brings about the change of state from a solid to a liquid is known as the *latent heat of fusion*. The heat required to convert a liquid to a gas is termed the *latent heat of evaporation* or *vaporization*. One of the interesting facts brought out previously is that a liquid, once brought to the boiling point, does not increase in temperature, but utilizes the heat energy it takes up in converting more liquid to a gas or vapor.

If we were creatures dwelling on another planet and accustomed to a normal atmospheric temperature of 250°F or 300°F, we would be able to utilize water to cool substances by immersing them in a bath of water. Since heat and solid are only relative to our senses, materials immersed in a pan of water will lose heat to it, and the heat extracted will be taken up by the water and used to generate steam. Through the extraction or flow of heat from objects so immersed, the water (assumed to be under the same normal sea-level pressures) would remain at 212°F, which would be cooler than the atmospheric temperature of our hypothetical planet. Fortunately for Earth dwellers, other liquids boil at temperatures much below that of water.

Refrigerants

Among common liquids that boil at a temperature below that of water is alcohol, with a boiling point of 173°F, and ether boils at ordinary summer heat, which is 94°F. Other substances boil at still lower temperatures. Carbon dioxide boils at –110°F, ammonia at –20°F, sulfur dioxide at 14°F, methyl chloride at –11.6°F, ethyl chloride at 54°F, Freon-12 at 21.6°, and Freon-22 at 41.4°F—all at atmospheric pressure encountered at sea level.

Those materials (solid or liquid) that vaporize or liquefy at comparatively low temperatures, and are suitable for use in refrigeration work, are termed *refrigerants*. The refrigerants are employed in specially designed apparatus so that the extraction of heat from rooms or perishables can be accomplished as inexpensively as possible.

The latent heat of vaporization of liquid refrigerants varies with the material and pressure at which vaporization is allowed to proceed. The latent heat of vaporization in Btu per pound of various refrigerants in common use and volatilizing at 0°F is presented as follows. There will be notes accompanying these refrigerants because some of them are now practically obsolete, but you may run across some older equipment that uses them:

- *Ammonia* (572.2)—Ammonia is considered a hazardous material, and you will encounter it only in very large commercial installations, where the hazards have been allowed for.
- *Methyl chloride* (176.0)—Methyl chloride mixed in the right proportions with air will burn or explode. It has not been used much in home refrigeration and is not being used today to any degree.
- *Ethyl chloride* (173.4)—Ethyl chloride has been practically eliminated from the refrigeration field.
- *Sulfur dioxide* (SO_2) (171.8)—Sulfur dioxide was once used to a large degree, but it is very toxic and the only places that you will encounter it are in very old domestic refrigerators.
- *Carbon dioxide* (CO_2) (117.5)—Carbon dioxide is not used in domestic refrigerators.
- *Freon-12* (71.0)—Freon-12 was very popular for a long time.
- *Freon-22* (100.7)—Freon-22 is no longer allowed to be used in domestic freezers.
- *Freon-502* (74.8)—Freon-502 is used in large air conditioning systems.
- *Freon-21* (119.4)—Freon-21 is used in industrial and commercial air conditioning.
- *R-134a*—R-134a is a new refrigerant used in automobile air conditioners.

From this list, it is observed that ammonia is an excellent refrigerant in that it absorbs a great quantity of heat as compared to other agents. However, it is toxic and classified as a hazardous material, so its use is confined mostly to domestic absorption-type refrigeration and large industrial manufacturers of ice. Methyl chloride is combustible under the right mixtures with air, and so it is seldom used any more. Sulfur dioxide (SO_2), although quite popular in the past, is extremely toxic and no longer used. It will be covered to some extent in this book, as you may come across an SO_2 system that needs repair. Carbon dioxide (CO_2) in solid form is used in transporting ice cream, fish, and similar products and where a great quantity of escaping gas fumes would not endanger human life. This substance is also known as dry ice.

It is very easy to understand that a vessel containing a refrigerant (which is allowed to absorb heat) will cause cooling or a refrigerating effect. Since heat (as energy) cannot be created or destroyed, it follows that heat removed from one body must show up in some form of energy in another body. Essentially, the removal of heat from one body and its transference to and dissipation by another body is refrigeration.

Use of Ice

Where ice is employed, the latent heat of fusion results in the cooling of other materials suffering such loss of heat energy that the heat is taken up by the ice and it reverts to water. The heat-laden liquid (water) is drained away, carrying its heat load out and away from the refrigerator. A graphical illustration of this cycle is shown in Figure 1-17.

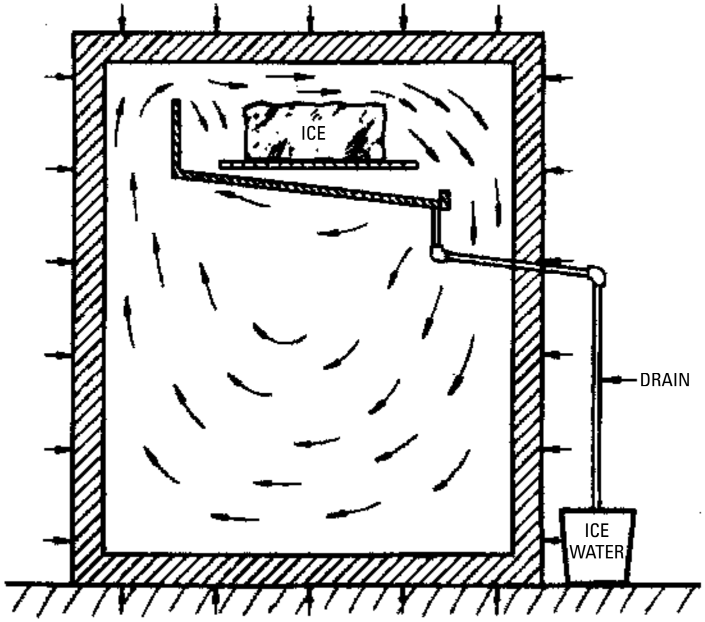


Figure 1-17 Example of the ice-refrigeration cycle.

Ton of Refrigeration

Because refrigeration was first produced by ice, the rate of removal of heat in a cooling operation is expressed in terms of pounds or tons of ice required per unit time, usually per day. It has been found that 1 lb of ice absorbs 144 Btu when it melts; 1 ton of ice consequently absorbs 2000 × 144, or 288,000 Btu. When 1 ton of ice melts in 24 hours, the rate is 288,000/24 or 12,000 Btu/hr, or 12,000/60 = 200 Btu/min. This rate has been officially designated as 1 ton of refrigeration and is the basis for rating all refrigerating machinery.

Evaporation

We previously discussed the quantities of heat required to bring about a change of state. Considerable heat energy is necessary when a change takes place from the liquid to the vapor phase. The term *evaporation* is well known to all. A common example of refrigeration by means of evaporation is found in the cooling of the human body by perspiration. The beads of moisture appear on the skin and, by evaporation into the air, actually maintain the body at a lower temperature than the surrounding atmosphere.

The normal internal temperature of human beings is 98.6°F. Everyone, at some time or another, has encountered atmospheric temperatures that were much higher than body heat, never realizing that it was evaporation that acted as the refrigerating element to keep the body cool. The human body is extremely sensitive and is always within a very limited range of temperature, allowing a variation of only a few degrees. For example, if the inside body temperature were raised above 108°F, death is almost certain; yet men have lived in atmospheres where temperatures were as high as 140°F and even higher. Their bodies were maintained at normal body heat by evaporation of moisture from their skins.

Another example of cooling by evaporation found in nature is the reduction in temperature that usually follows a summer rain. The required heat to bring about the evaporation of the rainwater is drawn from the atmosphere and the Earth and thus affects the coolness we note after such precipitation.

Pressure

In the previous instances of boiling temperatures, the wording *at pressures found at sea level* was used. For example, water boils at 212°F when it is heated in a vessel open to the atmosphere and when the experiment takes place at an altitude at or near sea level or under equivalent pressure. Regardless of how rapidly heat may be added to the open vessel of water, no rise in boiling temperature can be secured. The only result achieved is that steam is formed with greater rapidity commensurate with the amount of heat. If, however, pressure is allowed to enter as a factor in the conversion of a liquid to a vapor, or vice versa, it will be found that the boiling point is altered. We are all more or less familiar with the story of the man who tried to prepare his eggs on the top of a high mountain where atmospheric pressure because of altitude was so low that the water boiled at such a low temperature that it would not cook the eggs. From this simple example, it is apparent that the lower the pressure exerted on the liquid, the lower will be the boiling point.

By way of demonstrating the effect of pressure, let us take a sturdy flask, one capable of withstanding both heat and pressure, and pour in some water. We will assume it to be a glass vessel so that we can view the process and make certain when boiling actually occurs. A stopper provided with three perforations, so that a pressure gage, thermometer, and a valve can be inserted in the individual openings, is forced tightly into the neck of the flask. The completely assembled apparatus is then placed over a source of heat (such as a gas flame), as depicted in Figure 1-18. The exit valve is allowed to remain wide open. Soon the water will boil, and if the experiment is performed at or near sea level, the pressure gage will register 0 lb of pressure while the thermometer will indicate a temperature of 212°F. As long as the valve orifice is large enough to offer a free and unobstructed exit to all of the steam formed, no amount of heat will raise the temperature.

Now let us turn the valve so that the opening is slightly restricted. The pressure will build up, and by regulating the amount of gas burned and the exit valve so that a jet of steam is released, we can obtain or impose any pressure we desire on the vessel. Let us select a pressure of about 5 lb as the pressure desired. With proper regulation of the gas and steam valves, this value can easily be obtained and maintained. When secured, let us read the thermometer. By readjusting the steam and the gas valves, let us raise the pressure to 10.3 lb gage pressure (10.3 psig). At this point we would find the water would start to boil at 240.1°F. Again, we would find that no amount of heat added rapidly or otherwise to our flask would raise its temperature above that point, provided the pressure remained constant.

With heat added rapidly, more steam would be formed and we would, of course, be obliged to open the steam exit valve to maintain the pressure at the desired point. If only a little heat were added, we would have to close the valve somewhat to maintain the pressure. By referring to a table giving the properties of saturated steam, various

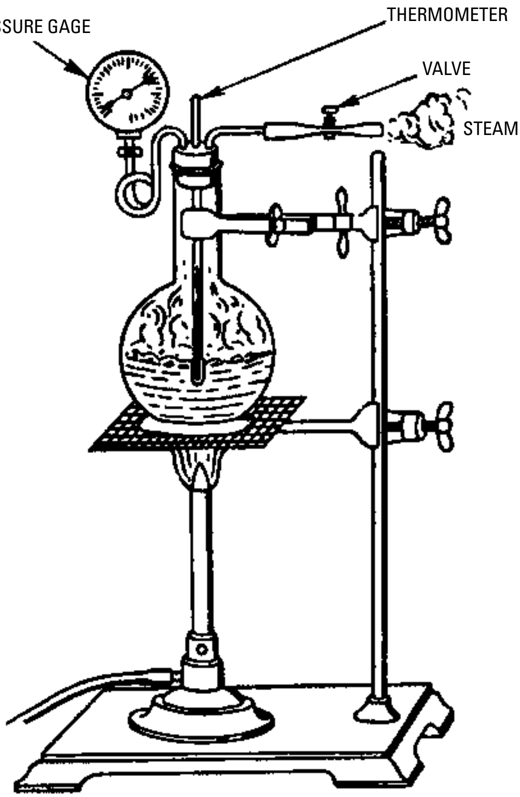


Figure 1-18 Apparatus needed for evolution of pressure-temperature relations of water.

gage pressures and corresponding temperatures can be determined. For example, if our boiler were of heavy steel, we could increase the pressure to a point where the gage would indicate 100.3 lb, at which point the boiling of the water would occur at 338.1°F. With the increase in pressure, the boiling point advances, and, again, in accordance with this pressure-temperature relationship, the boiling point is lowered as the pressure is reduced.

When the pressure decreases, the boiling temperature of a liquid will decrease also. For example, water at an atmospheric pressure equal to 4800 feet above sea level boils at approximately 202°F. Water at 10,000 feet altitude boils at approximately 193.2°F.

Pressure-Temperature Relations

Most bodies expand when heated, whether they are of solid, liquid, or gaseous form. If the expansion is restricted, great forces are set up by the body in the effort to expand. Every gas or vapor confined in a closed vessel exerts a certain pressure against the restraining walls of the vessel. The pressure imposed depends on the amount and temperature of the gas. The natural tendency for a confined gas is to expand when the temperature is increased. If the gas is contained in a vessel with rigid walls, the pressure will increase. This is of such importance that a law has been formulated to express it, stating that the pressure exerted by a gas or vapor in a closed vessel is *directly proportional to the absolute temperature of the gas*.

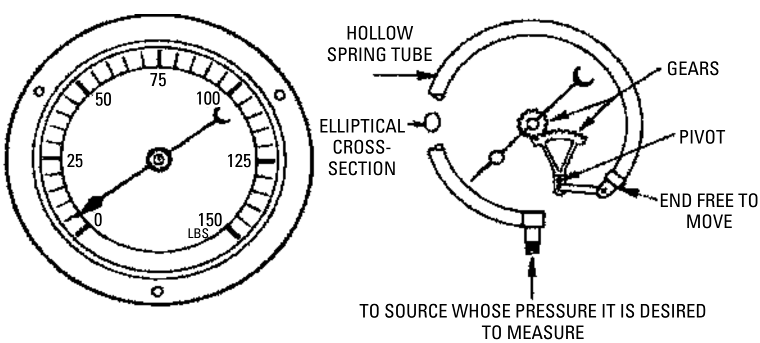
Low Pressures

Before proceeding with our experiments with lower pressures, it is best that we understand the comparative terms. As pointed out previously, objects at sea level are taken as being subjected to a certain constant pressure. This pressure is caused by the weight of the atmosphere, which at sea level is found to exert a pressure of 14.7 psi of surface.

The atmosphere is like a high sea. If we dive into water, similar conditions are encountered because the deeper we go, the greater the pressure becomes because of the volume above. As we rise, the pressure becomes less. Modern aircraft designers take into consideration the rarefication of the atmosphere with increasing altitude and know that at a certain height the air density will not allow an airplane to rise any higher. This they term the *ceiling,* or *density,* at which further altitude for each particular design is not possible. Since pressure is so instrumental in a great many engineering calculations, a universal standard was necessary. Thus, the pressure existing at sea level was accepted as a standard, being termed 0 lb gage, or 14.7 lb absolute.

In the manufacture of gages, the instruments are usually set up to indicate zero pressure at normal atmospheric sea-level density. It must be remembered that this setting at zero is made when the actual and true pressure is in reality 14.7 psi, the pressure that the weight of the atmosphere exerts at sea level. Nearly all steam and pressure gages are set in this fashion, and the pressures indicated by these instruments are termed *gage pressures*. The usual increments are in 1-, 5-, and 10-lb readings, so that only approximate pressures can be determined. For most commercial applications, these roughly calibrated gages are sufficiently accurate and quite inexpensive.

For fine work and experimental and scientific determinations, another pressure scale is employed. This scale is called the *absolute pressure* and is based on true pressures. Zero on this scale indicates no pressure at all or, in other words, a perfect vacuum. Figure 1-19 shows an external view and construction details of a typical pressure gage. It consists of a metal tube of elliptical cross section bent into a nearly complete ring and closed at one end. The flatter sides of the tube form the inner and outer sides of the ring. The open end of the tube is connected to the pipe through which the liquid under pressure is admitted. The closed end of the tube is free to move. As the pressure increases, the tube tends to straighten out, moving a pointer through a lever-and-gear arrangement. The scale is graduated directly in psi.



(A) Front view. (B) Internal connection.

Figure 1-19 External view and construction details of a typical pressure gage.

On the pressure-gage scale, the normal atmospheric pressure found at sea level (14.7 psi) is indicated as 0 lb pressure, and below that point the term of pressure is converted to vacuum, or inches of mercury (in Hg). A long glass tube, sealed at its upper extremity, is exhausted of air and filled with mercury, as shown in Figure 1-20, its lower open end dipping into a reservoir of mercury open to the atmosphere or to the pressure imposed. Such a device is called a *barometer* and is used to measure the pressure in terms of inches of mercury in place of a term such as pounds under zero gage pressure. Just as variance in temperature will cause a rise or fall of the mercury level in a thermometer, so will a difference of pressure exerted on the mercury in the reservoir of the barometer bring about

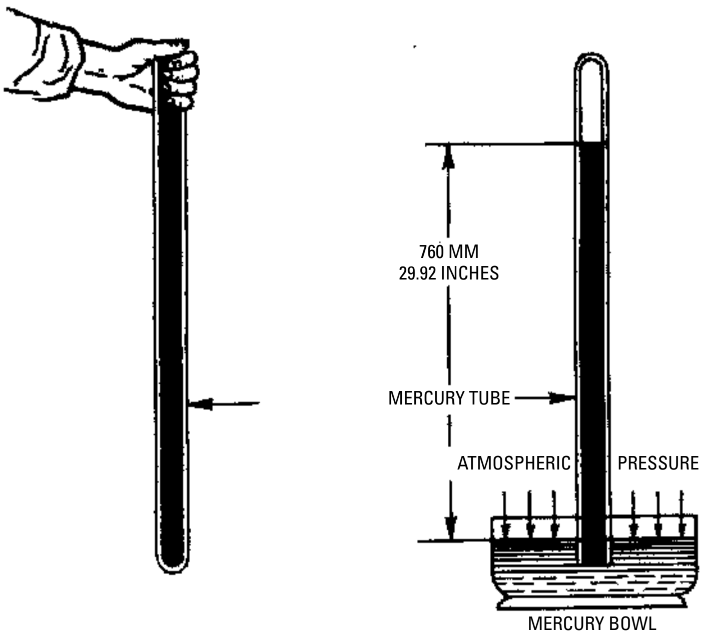


Figure 1-20 Balancing the pressure of air with a mercury column.

a certain reading in terms of inches of mercury. The pressure of the atmosphere or any other pressure exerting its force on the mercury exposed in the reservoir drives the liquid up into the evacuated end to a height corresponding with the pressure. In a way, the barometer is much like the thermometer because it indicates a value by the height of a mercury column, one recording in terms of temperature and the other in terms of an equivalent of pressure.

On the absolute scale, it will be observed that pressures begin from an absolute zero or perfect vacuum and that a gage calibrated in this manner will indicate 14.7 lb at normal sea-level pressures. On the other hand, a gage calibrated in the fashion employed for the commercial field of refrigeration and air conditioning will indicate this latter pressure as 0 lb where normal atmospheric conditions prevail.

Effect of Low Pressures

In the previous experiments, we determined that whenever pressure was increased, the boiling point was raised. Making use of the same apparatus, let us investigate the effect of a reduced pressure on the evaporating temperature. To do so, we will again place a quantity of water in the flask, force the stopper containing a pressure gage, thermometer, and control valve tightly in place so that no leakage can occur to mar the accuracy of our experiment, and place the apparatus over a gas flame. This time, however, instead of using the exit valve for holding the steam within the vessel to cause a pressure to build up, we will connect a vacuum line to it and use it to prevent the steam from being drawn out or evacuated so that a certain low pressure (or inches of vacuum) can be maintained within the vessel.

Referring to the pressure scale (Figure 1-20), we find that 30 in Hg is roughly equivalent to 15 lb on the gage-pressure scale, so that a 2-inch vacuum means a pressure of 1 lb less than the sea-level atmospheric pressure. For our first experiment, let us take a pressure of 14 lb absolute, or 0.7 lb less pressure than normally exerted. From a table on the properties of saturated steam, we will find that this corresponds to 1.42 in Hg on the gage-pressure scale.

By regulating the exit valve so that the proper vacuum or low pressure is obtained, we can easily arrive at a point where boiling occurs at the constant pressure we desire. At this point of 14 lb absolute pressure, we would find that the water boils at 209.55°F. Then, we can open the exit valve a trifle more and regulate it so that a pressure of 12 lb absolute (5.49 in Hg) is maintained. This pressure results in a boiling point of 201.96°F.

The fact is thus brought out that the pressure to which water is subjected has just as much to do with its boiling as does the temperature. This holds true for any liquid because evaporation can be made to occur at any temperature above its freezing point if the pressure to which it is subjected is made low enough.

As you progress in your studies, you will find that sometimes you will have to put a vacuum pump on a refrigeration system to get rid of all unwanted gases and water. At sea level, theoretically you should get the vacuum pump to pull down to 29.92 in Hg. This will not be possible because the pump you will be using will not pull a perfect vacuum. Depending on the atmospheric pressure and efficiency of your pump, you will probably be able to pump down to between 27 and 28 in Hg. At a 5000-foot altitude you may reach approximately 23 inches of vacuum.

Heat Transmission

In refrigeration, we are interested in getting the heat contained in a room or refrigerator to a medium that will effect its removal. Ice is a simple method and formerly was widely used. The more modern refrigerator is equipped with a cooling apparatus that supplants ice and provides for more constant and cooler temperatures. Before we can study how heat is taken up and removed mechanically, it is imperative to learn the behavior of heat.

One of the most important laws has already been mentioned—the flow of heat from a body of higher temperature to one having a lower sensible heat. Never of its own accord will heat or water flow uphill or in the opposite direction. Therefore, it follows that heat in a refrigerator or room will flow to the cooler object (such as ice or the cooling device). The transmission of heat may be accomplished in three ways: by *conduction*, *convection*, and *radiation*.

Conduction

Conduction is the transference of heat by molecular impact from one particle to another in contact. For example, if the end of a bar of iron is heated in the fire, some of the heat will pass through the bar to the cooler portion. Heat traveling in a body or from one body to another where the two are in intimate contact is termed *conduction*, as shown in Figure 1-21.

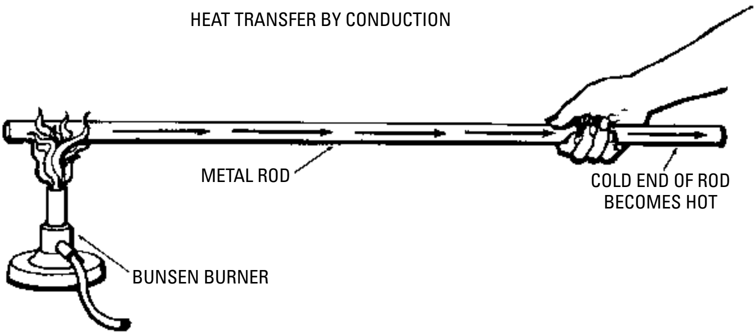


Figure 1-21 Transfer of heat by conduction.

Metals are usually splendid heat conductors. Every material has a conduction value—some good (like the metals), others mediocre, and a few very poor. For example, heat will quickly pass through a piece of copper but will have considerable difficulty in passing through a piece of cork. The materials that have very low heat conductivities are termed *heat insulators*. Even the very poorest conductors (or insulation materials) allow a certain amount of heat to pass through. No material exists that offers a perfect barrier or resistance to the passage of heat.

Convection

Convection is the principle used in hot-air heating. Air that is free to circulate (such as in any air body of appreciable size) will be set into motion where a difference of temperature occurs because it will absorb heat from the warmer wall, become heated, expand, and become lighter. The heated portion of the air will rise, and cooler air will move into its place, which, in turn, will become heated. The heated portion of the air eventually moves over to the colder wall, and the heat flows from the air to the colder object. Thus, any body of air capable of motion will transmit heat by convection. Hot-air and hot-water heating systems work on the convection principle. They convey heat by bodily moving the heated substance from one place to another, as shown in Figure 1-22.

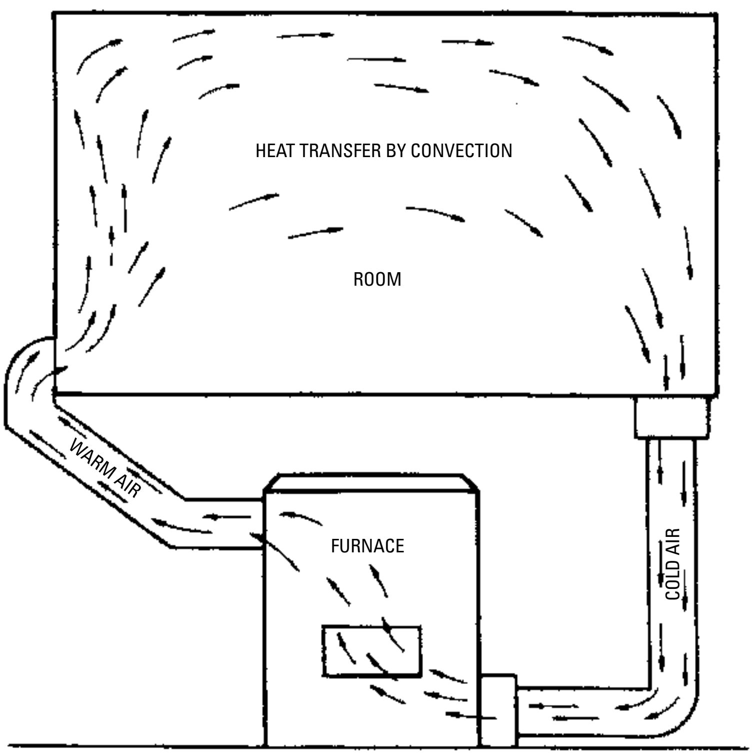


Figure 1-22 Transfer of heat by convection.

The most efficient heat insulation known is a vacuum; but except for very small containers, it is structurally impractical to employ it commercially. The next best insulating medium is air subdivided into the smallest possible units so that it is still, or stagnant. Air that is contained in spaces of appreciable size (such as between the double walls of refrigerators) will circulate and transmit heat by convection. Cork is an insulation material of a high order and has great resistance to the passage or transmission of heat because of its air content. The air cells in cork are of such minute size that the air trapped in them is so restricted that only a little circulation is possible. For all practical purposes, it is still air, and little or no convection occurs.

Radiation

Heat energy transmitted through the air in the same way light is sent out by a lighted lamp, a radiant heater, or the sun is called *radiated energy,* as shown in Figure 1-23. Large cold-storage warehouses, auditoriums, theaters, and homes are built with consideration of the heat evolved through radiant energy of the sun. Small household appliances rarely require the consideration of any radiant-heat factor because they are used in existing structures without any change in building design and are sheltered from direct heat.

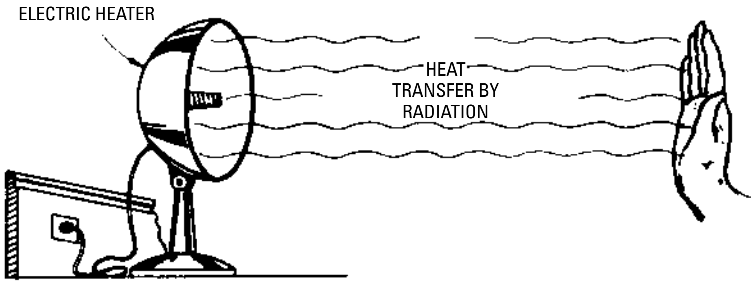


Figure 1-23 Transfer of heat by radiation.

Condensation

In a previous experiment we illustrated that a liquid, heated to the boiling point corresponding to the pressure imposed, will assimilate heat and produce a vapor, or gas. The heat taken up by the liquid is used to speed up molecular activity until a vapor is evolved. It stands to reason, then, that any vapor, or gas, contains a considerable quantity of heat.

In accordance with the foregoing experiments, we found that heat itself always flows from the warmer to the colder body. To prove this again and illustrate just what *condensation* is, let us take a vessel,

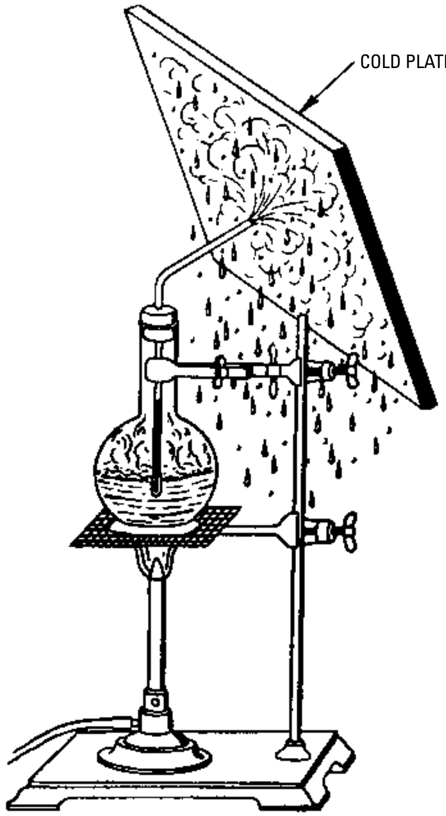


Figure 1-24 Example of condensation.

fill it about half full of water, and bring it to the boiling point over a gas flame, as shown in Figure 1-24. As soon as steam is generated, let us take a dry, cold plate and hold it at an angle over the jet of steam as it issues from our crude boiler. It will be observed that the steam impinges on the plate. There it is converted again into water, the droplets forming and dripping off the edge of the plate. This was caused by the hot steam giving up its heat load (latent heat of vaporization) to the cold plate, the heat flowing from the steam to the plate.

To make sure that this is really the case, let us take the vessel we used in a previous experiment and again fill it with water, secure the stopper in place, and apply heat. This time let us set the exhaust valve so that we can maintain a few pounds of steam pressure (perhaps 5 lb). While steam is generating and we are regulating the valve to get this constant pressure, let us build a condenser. This is easily done; all we need is about 12 feet of copper tubing. Any tubing or material will do, but copper is so easily bent and formed that it is not a difficult task to wind it in a spiral form, such as shown in Figure 1-25.

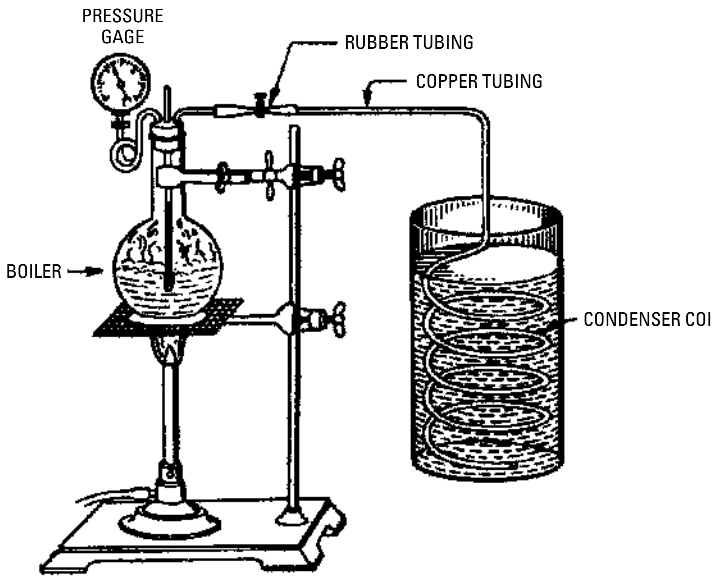


Figure 1-25 Example of condensation by steam vapors.

The formed coil is submerged in a pail of cold water with its open end at the bottom. The top end is supplied with a piece of rubber tubing. When the gas and steam valves have been regulated so that there is a constant steam pressure of 5 lb within the boiler and a fair amount of steam being exhausted from the exit, quickly connect the rubber tube on the condenser to the steam pipe. It will be observed that the pressure gage has dropped somewhat. After the steam has blown out the air contained in our condenser, we will see that no steam is issuing from the open end of the copper coil submerged in the water.

If we weighed the steam-generating apparatus and the condenser equipment before and after the experiment, we would find that the boiler lost weight through evaporation of water, whereas the condenser gained through condensation exactly what had been lost. If a thermometer had been placed in the condenser water and a reading taken at the start and end of our experiment, it would have been found that the water increased in temperature. In fact, if we ignore various losses, the amount of steam condensed can be estimated from the increase in temperature of the water.

It is apparent that the steam is easily convertible to its liquid form if the heat of vaporization is extracted. Then, too, our old rule would apply, not only to steam but also to all vapors and gases. That is, if we increase the pressure, we will find that the gas will condense at the lower temperature found in the condenser because the temperature of the gas will be raised far above the condenser temperature. Thus, heat will flow from the hot gas to the cold condenser. The greater the temperature difference, the faster will be the heat exchange.

In refrigeration and air conditioning applications, the gases employed as a refrigerating medium, for any given pressure, have a corresponding temperature at which they condense, or liquefy. Where both gas and liquid are present in the same vessel and the closed container is heated to cause boiling, the temperature of both gas and liquid will be the same at the boiling point. If pressure is increased, the boiling point will be raised. Above a certain point, the gas will cease to have any latent heat of vaporization, and it will remain a gas regardless of the intensity of the pressure imposed.

Altitudes and Refrigeration Capacity

As altitude increases, there will be a decrease in the efficiency of refrigeration because of the lower air pressures. Thus, the cooling power of the condenser of the refrigerator and the horsepower of the motor naturally decrease. Do not be alarmed about this on home refrigerators. They are oversized for sea level and work well at higher altitudes but with less efficiency.

Atmospheric Factors in Refrigeration

Dry-bulb temperature is the temperature that the plain thermometer will register. *Wet-bulb temperature* is that temperature observed if an ordinary thermometer bulb is covered with material such as linen, wet with distilled water, and exposed to atmospheric evaporation. Table 1-11 shows dry- and wet-bulb temperatures as related to summer air conditioning design conditions.

The amount of refrigeration required is affected not only by dry-bulb temperature but also by humidity, indicated by wet-bulb temperature, because of the latent heat required to condense the water

Table 1-11 Typical Outdoor Design Conditions for Air Conditioning

City	*Dry-Bulb Temp °F*	*Wet-Bulb Temp °F*
Birmingham, Alabama	95	78
Phoenix, Arizona	105	76
Los Angeles, California	90	70
Denver, Colorado	95	64
Washington, D.C.	95	78
Tampa, Florida	95	78
Atlanta, Georgia	95	76
Chicago, Illinois	95	75
New Orleans, Louisiana	95	80
Boston, Massachusetts	92	75
St. Louis, Missouri	95	78
New York City, New York	95	75
Cincinnati, Ohio	95	78
Dallas, Texas	100	78
Seattle, Washington	85	65

vapor, which will collect as frost on the evaporator of the air conditioner.

Refrigerant pressure versus temperature is valuable information to have when repairing or installing refrigeration equipment. Table 1-12 is based on gage pressure in psi at sea level and may be used in home refrigeration and air conditioning to determine the relative temperatures of evaporators. In commercial refrigeration (where pressure switches instead of thermostats are used to control the temperature desired), the table can be an extremely valuable tool in setting the desired temperatures. On present-day home refrigeration and air conditioning, R-12 and R-22 are the most commonly used refrigerants. In this text, you will notice other refrigerants have been included, as you no doubt will come across some of these in older machines that you may be called upon to service.

Refrigeration by Vaporization

The temperature at which a liquid boils or vaporizes is called its boiling point. For liquids with very low boiling points, it is not necessary to supply heat by fire or other artificial means because the heat in surrounding objects may be sufficient to cause boiling or

Table 1-12 Refrigerant Pressure vs. Temperature*

	Refrigerant			
Temp.,°F	*R-12*	*R-22*	*R-502*	*R-717 NH3*
−40	11.0†	0.5	4.1	8.7†
−35	8.4†	2.6	6.5	5.4†
−30	5.5†	4.9	9.2	1.6†
−25	2.3†	7.4	12.1	1.3
−20	0.6†	10.1	15.3	3.6
−15	2.4	13.2	18.8	6.2
−10	4.5	16.5	22.6	9.0
−5	6.7	20.1	26.7	12.2
0	9.2	24.0	31.1	15.7
5	11.8	28.2	35.9	19.6
10	14.6	32.8	41.0	23.8
15	17.7	37.7	46.5	28.4
20	21.0	43.0	52.5	33.5
25	24.6	48.8	58.8	39.0
30	28.5	54.9	65.6	45.0
35	32.6	61.5	72.8	51.6
40	37.0	68.5	80.5	58.6
45	41.7	76.0	88.7	66.3
50	46.7	84.0	97.4	74.5
55	52.0	92.6	106.6	83.4
60	57.7	101.6	116.4	92.9
65	63.8	111.2	126.7	103.1
70	70.2	121.4	137.6	114.1
75	77.0	132.2	149.1	125.8
80	84.2	143.6	161.2	138.3
85	91.8	155.7	174.0	151.7
90	99.8	168.4	187.4	165.9
95	108.2	181.8	201.4	181.1
100	117.2	195.9	216.2	197.2
105	126.6	210.8	231.7	214.2
110	136.4	226.4	247.9	232.3
115	146.8	242.7	264.9	251.5
120	157.6	259.9	282.7	271.7
125	169.1	277.9	301.4	293.1
130	181.0	296.8	320.8	
135	193.5	316.6	341.3	
140	206.6	337.2	362.6	

*Gage pressure in psi at sea level.
†In Hg below standard atmosphere.

vaporization. This is true, for example, with anhydrous ammonia, the boiling point of which is –28°F, which is sufficiently low to cause it to boil violently when placed in an open vessel at ordinary temperatures. The absorption of heat by the vaporization of the ammonia will cause the outside of the container to become heavily frosted by moisture condensed and frozen from the air immediately surrounding the container.

Since the boiling temperature of any liquid may be changed by the pressure exerted upon it, it is easy to cause a liquid refrigerant to boil at any desired temperature by placing it in a vessel where the required pressure is maintained. The process of boiling or vaporization by which a liquid is changed to a vapor can be reversed. That is, the vapor can be reconverted into a liquid by the removal of heat. This is called *condensation*. An increase in pressure (by raising the boiling point) will assist in the condensation of the vapor. Liquids used as refrigerants must be recovered because of their initial cost, and the process of condensation is usually employed for this purpose.

Although some liquids boil at temperatures suitable for refrigeration, comparatively few possess all the requirements of a practical refrigerant. Those used commercially are Freon-12 (R-12), Freon-22 (R-22), Freon-502 (R-502), and ammonia (R-717, NH_3). Others used previously are sulfur dioxide, methyl chloride, and carbon dioxide. It is very questionable that you will run into any of the latter refrigerants now. For large installations, anhydrous ammonia will be used as well as lithium bromide. The first four refrigerants mentioned, listed in Table 1-12, are the refrigerants that you will encounter today in household and industrial refrigeration.

Basic Systems

Refrigeration is distributed by several methods. In household and commercial refrigeration, the one you will find in most cases is the *direct expansion system*. The volatile refrigerant is allowed to expand in a pipe placed in the room or household refrigerator to be cooled, where the refrigerant absorbs its latent heat of evaporation from the material to be cooled. This method is used in small cold-storage rooms, constant-temperature rooms, freezer rooms, household refrigerators, and where possible losses caused by the leakage of refrigerant would be low.

In the *indirect system,* a refrigerant medium (such as brine) is cooled down by the direct expansion of the refrigerant and is then pumped through the material or space to be cooled, where it absorbs its sensible heat. Brine systems are used to advantage in large installations where the danger of the large amount of refrigerant is important and the rooms or series of rooms have fluctuating temperatures. They will also be found in homes that are heated with hot-water systems, except that instead of brine ordinary water is used with temperature controls for keeping the water from freezing. You will find these systems primarily used in air conditioning.

For small refrigerated rooms or spaces where it is desired to operate the refrigerating machine only part of each day, the brine coils are supplemented by holdover (or congealing) tanks. A *holdover tank* is a steel tank containing strong brine in which direct expansion coils are immersed. During the period of operation of the refrigeration machine, the brine is cooled down and is capable of absorbing heat during the shutdown period, the amount depending upon the quantity of brine, its specific heat, and the temperature head. *Congealing tanks* serve the same general purpose, but operate on a different principle. Instead of strong brine, they contain a comparatively weak brine solution, which freezes or congeals to a slushy mass of crystals during the period of operation. In addition to its sensible heat, this mass of congealed brine is capable of absorbing heat equivalent to its latent heat of fusion.

Laws of Gases

Extensive investigations of the behavior of gases have shown that a given weight of gas expands or contracts uniformly $^1/_{459}$ of its original volume for each degree it is raised or lowered above or below 0°F, provided the pressure on the gas remains constant. This fact is known as the *law of Charles*. Following this same reasoning, we find that at −459.62°F, a gas would cease to exist. This assumption, therefore, establishes −459.62°F as *absolute zero*. Actually, this temperature or condition has never been attained. The *law of conservation of matter* states that matter can be changed from one form into another. Temperatures within a few degrees of absolute zero have been reached when liquefying oxygen, nitrogen, and hydrogen, but these (like other gases) change their physical state from gas into liquid and fail to disappear entirely at these low temperatures.

The fact that absolute zero has never been reached is also explained by another law, known as the *law of conservation of energy*. It has already been explained that heat is a form of energy. This law is stated as follows: Energy can be neither created nor destroyed, though it can be changed from one form into another. Table 1-13 gives factors for converting energy from one form to another.

Having considered the effect of temperature on a gas, the effect of pressure on gases to aid the study of refrigeration must be considered next. In 1662, Robert Boyle announced a simple relationship

Table 1-13 Energy Conversion Factors

Unit	*Equivalent*
1 watt hour (Wh)	3.411 British thermal units (Btu)
1 British thermal unit (Btu)	0.252 calorie (cal)
1 calorie (cal)	3.968 British thermal units (Btu)
1 pound melting ice equivalent (MIE)	144 British thermal units (Btu)
1 British thermal unit (Btu)	0.00695 pound melting ice equivalent (MIE)
1 pound melting ice equivalent (MIE)	36.3 calories (cal)
1 calorie (cal)	0.0276 melting ice equivalent (MIE)

existing between the volume of a gas and the pressure applied to it, which has since become known to scientists as *Boyle's law* and may be stated as follows: *At a constant temperature, the volume of a given weight of gas varies inversely as the pressure to which it is subjected*. The more pressure applied to a gas, the smaller its volume becomes if the temperature remains the same; likewise, if the pressure is released or reduced, the volume of the gas increases. Mathematically, this is expressed as follows:

$$PV = p \times v$$

where the following is true:

P = pressure on the gas at volume V

p = pressure on the same weight of gas at volume v

Boyle's law has been found to be only approximately true, especially for refrigerant gases, which are more easily liquefied. The variations from the law are greater approaching the point of liquefaction, or condensing point, of any gas, although the material movement of air is determined by this law.

It will be found that if the temperature is held constant and sufficient pressure is applied to a given weight of gas, it will change from the gaseous state into the liquid state. The point at which this change of state takes place is known as the *point of liquefaction* or *condensing point*.

It should now be evident that a definite relationship exists between the pressures, temperatures, and volumes at which a given weight of gas may exist. This relationship is used extensively in scientific work. It is known as the *combined law of Boyle and Charles* and may be expressed mathematically as follows:

$$\frac{PV}{T} = \frac{pv}{t}$$

where pressures P and p are expressed in the absolute pressure scale in pounds per square foot, volumes V and v are expressed in cubic feet, and temperatures T and t are expressed in degrees on the absolute temperature scale.

When the pressure, temperature, or volume of a gas is varied, a new set of conditions is created under which a given weight of gas exists in accordance with the preceding mathematical equation. If a gas is raised to a certain temperature (which varies with each individual gas, no matter how much pressure is applied to it), it will be found impossible to condense it. This temperature is known as the *critical temperature*. The pressure corresponding to the critical temperature is termed the *critical pressure*. Above the critical points it is impossible to vaporize or condense a substance.

When a liquid is evaporated to a gas, the change of physical state is always accompanied by the absorption of heat. Evaporation has a cooling effect on the surroundings of the liquid since the liquid obtains the necessary heat from its surroundings to change the molecular structure. This action takes place in the evaporator of a refrigeration system. Any liquid tends to saturate the surrounding space with its vapor. This property of liquids is an important element in all refrigeration work.

On the other hand, when a gas is condensed into a liquid, the change of physical state is always accompanied by the giving up of heat. This action takes place in the condenser or the condensing unit of the refrigeration system because of the mechanical work exerted on the gas by the compressor. If gas or liquid is placed in a closed container and the temperature of the container changed, it will be found that the pressure exerted by that gas in the container is directly proportional to the absolute temperature. Thus, if the temperature is raised, the pressure increases; if the temperature is lowered, the pressure decreases.

The Refrigeration Cycle

The refrigeration cycle is simply a means of heat extraction. In a compression system four distinct parts are required:

- Compressor (pump)
- Condenser-receiver
- Evaporator (cooling coils)
- Expansion valve (pressure-reducing device)

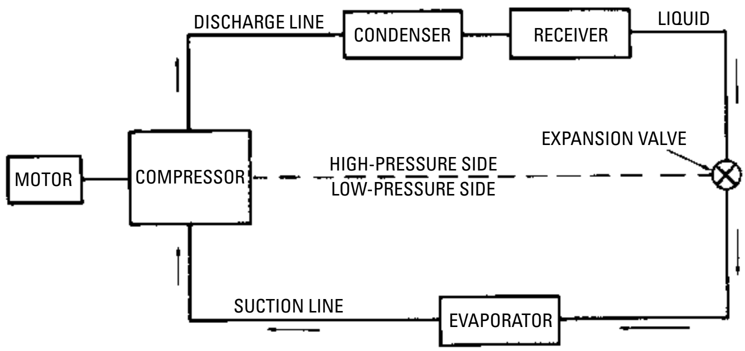


Figure 1-26 Basic refrigeration circuit.

Figure 1-26 shows schematically the parts that make up the compression system. In this process, a refrigerant is used that can be alternately vaporized and liquefied. The heat energy required to change the liquid refrigerant into a gas is obtained from the air space surrounding the evaporator. This low pressure is then drawn through the suction line and into the compressor. In passing through the compressor, the heat-laden gas is raised from the low pressure in the suction line to a higher pressure, thereby raising its temperature. It is then forced from the compressor through the discharge line into the condenser, where the heat is removed from the vapor by means of natural air circulation. The removal of heat from the vapor causes it to liquefy and then flow from the bottom of the condenser and into the receiver.

The expansion valve is adjusted to control the flow of refrigerant into the evaporator at a rate that is sufficient to maintain a desired temperature. In this connection, it should be noted that once the refrigerant has returned to a liquid state, it is again ready to be admitted through the expansion valve (or other pressure-reducing device) to the evaporator. In the evaporator, the pressure is reduced, the boiling point is lowered, and vaporization takes place, resulting in extraction of heat. This action is repeated continually as long as the compressor is running.

Pressure

It is of the utmost importance that the student of refrigeration understands the meaning of the various kinds of pressure as related to refrigeration.

Atmospheric Pressure

Atmospheric pressure is pressure that is exerted by the atmosphere in all directions, as indicated by a barometer. Standard atmospheric pressure is considered to be 14.695 psi (usually written 14.7 psi), which is equivalent to 29.92 in Hg.

Absolute Pressure

Absolute pressure is the sum, at any particular time, of gage pressure and atmospheric pressure. Thus, for example, if the pressure gage at one particular time reads 53.7 lb, the absolute pressure will be $53.7 + 14.7$, or 68.4 psi. The preceding definitions may be written as follows:

$$\text{Absolute pressure} = \text{gage pressure} + 14.7$$

$$\text{Gage pressure} = \text{absolute pressure} - 14.7$$

where 14.7 is the normal atmospheric pressure.

Summary

In this chapter, the fundamental principles of refrigeration have been presented. Since refrigeration deals with the removal of heat from space or material substances, it is important that the student or service technician clearly understand the method used in refrigeration, as well as the nature of heat, heat of vaporization, latent heat, and change of state.

It should be noted that modern refrigeration is accomplished simply by a change of state from liquid to gas in various types of refrigerants, the change of its physical state being accomplished by the absorption of heat. The laws of gases have been fully treated in addition to the energy sources utilized for pumping the refrigerant through the refrigeration system components.

The measurement of heat and a simple calculation for removal of heat from various material substances using specific heats and Btu values will further assist in understanding the nature of refrigeration, its laws, and finally, its utilization.

Review Questions

1. What is meant by units of heat?
2. What is meant by conduction, convection, and radiation?
3. How does specific heat affect various substances?
4. How does latent heat affect the change of state in various fluids?

5. What is meant by heat of vaporization?
6. Name various commercial refrigerants.
7. How is refrigeration accomplished?
8. What is meant by an indirect system of refrigeration?
9. Define the law of conservation of matter.
10. What is the absolute zero reading on the Fahrenheit scale?
11. How many calories are represented by 1 Btu?
12. Define Boyle's law.
13. State the law of conservation of energy.
14. State the relation between pressure, temperature, and volume in a given weight of gas.
15. Explain how the refrigeration cycle is accomplished.
16. What will be the temperature reading on the Fahrenheit scale when the Celsius thermometer reads $-10°$?
17. Give the relation between absolute and gage pressure.
18. Why was it necessary to develop new refrigerants?
19. What is the condition of an object at absolute zero?
20. Give four qualities of steam that make it like all trite vapors or gases.
21. What is the difference between a square foot and a cubic foot?
22. What is a slug?
23. What is the difference between a short ton and a long ton?
24. What is specific gravity?
25. What is the atmospheric pressure? Why is it important to know when working with refrigerants?
26. How is Boyle's law used in reference to refrigerants?
27. What relationships does Charles' law deal with?
28. What is isothermal change?
29. How do you convert °C to °F?
30. How do you convert °F to °C?
31. What is the difference between kinetic and potential energy?