



AP* PHYSICS B

THERMAL PHYSICS: HEAT & THERMODYNAMICS

What you already know from Chemistry about temperature and heat:

- Never, ever forget that **temperature** is the average kinetic energy of molecules ($K_{\text{avg}} = \frac{1}{2}mv^2$)
- Metals conduct heat and electricity due to mobile, delocalized electrons.
- **Thermal contact** (commence heat exchange) leads to **thermal equilibrium** (cease heat exchange)
- Heat moves along a gradient from hot to cold until equilibrium is established
- **HEAT**--energy exchanged between objects because of a temperature difference
- A thermometer reaches thermal equilibrium once it stabilizes with the substance being measured

There are **FOUR Laws of Thermodynamics** that are inconveniently numbered!

- **Zeroth Law of Thermodynamics**--If two substances are in thermal equilibrium with a third system, then they are in thermal equilibrium with each other. What a silly name! Why do they call it that? They had already named the other laws and discovered this obvious one still needed stating, so they had to go backwards to zero.
- **First Law of Thermodynamics**--most commonly known as The Law of Conservation of Energy which *really* means that the total Energy, U of the universe remains constant!
- **Second Law of Thermodynamics**--The entropy [S --I know, the word entropy has no "S" in it--I don't decide these things! It's just a fancy word for disorder or chaos.] of the universe is always increasing. [Think about all the energy it takes to keep your room clean and how it always gets messy "by itself".]
- **Third Law of Thermodynamics**--The entropy of a pure crystalline substance at 0 K is equal to zero. [Perfect order and no molecular motion \therefore no disorder or chaos. Not actually attainable since electrons are still moving.]

THERMOMETERS AND TEMPERATURE SCALES

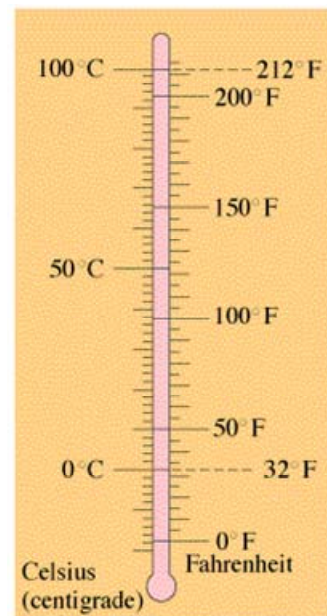
How thermometers work depends on the following **physical** properties:

- Δ Volume of a liquid [Δ is the capital Greek letter delta--it's used to mean **change** which is generally computed "final minus initial"]
- Δ Length of a solid
- Δ Pressure of a gas held at constant volume
- Δ Electric resistance of a conductor (that's how the CBL probe works--very accurate!)
- Δ Color of a very hot object such as red hot iron, sun, stars

Mercury or Alcohol Thermometers

A glass tube encasing a capillary tube filled with either of these liquids-- as the liquid is warmed it expands and has nowhere to go but up! It contracts upon cooling.

- $\Delta \text{Temperature} \propto \Delta \text{length of a column of liquid}$
(\propto means proportional and is often replaced with an equal sign accompanied by a constant with goofy units!)
- Calibrated using the freezing point, FP [0 °C, 32 °F, 273.15 K] of water and the boiling point, BP of water [100 °C, 212 °F, 373.15 K]. Celsius and Kelvin are broken into 100 equal segments between calibration points while Fahrenheit is broken up into 180 segments. Why is the Fahrenheit scale so odd? Daniel Fahrenheit developed a temperature scale that was also broken into 100 units, but he set 0 °F equal to the coldest temperature he could achieve using a sample of salt, ice and water and 100 °F as his own body temperature (very portable, but perhaps unreliable!).

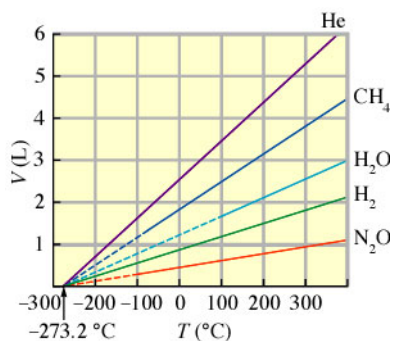


Two disadvantages:

- Hg and Alcohol have 2 different expansion coefficients [expand by different amounts for each degree they are warmed]. This means one may read 59.0 °C while the other reads 59.6 °C when immersed in the *same* substance. That's trouble!
- Limited temperature range--Hg freezes at -39 °C and both liquids vaporize.

Constant-Volume Gas Thermometers and the Kelvin Scale

- When the bulb is immersed the gas expands or contracts. The tube on the right is raised or lowered so that the gas is always experiencing the SAME pressure on the left side of the apparatus. The height change h is THEN read on the right side of the apparatus.



- Temperature readings are nearly independent of the gas used, unlike Hg or alcohol.
- Notice all gases extrapolate to -273.15 °C when pressure goes to zero! This was inspiration for the Kelvin scale where a degree C = a *kelvin* of **change**

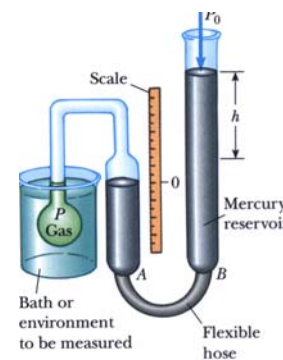
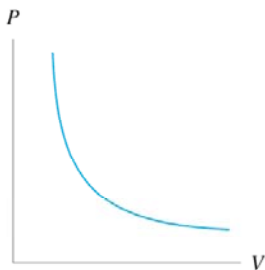
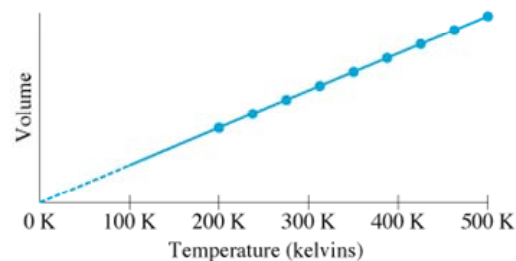
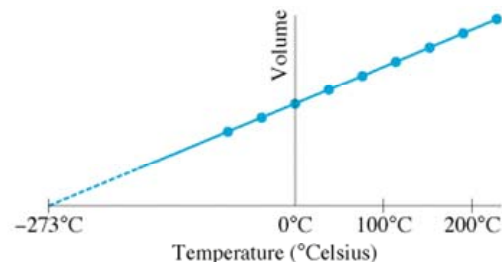


FIGURE 10.2
A constant-volume gas thermometer measures the pressure of the gas contained in the flask on the left. The volume of gas in the flask is kept constant by raising or lowering the column of mercury such that the mercury level remains constant.



So, 0 K = absolute zero \therefore all molecular motion stops \therefore no K_{avg}
 \therefore no Pressure (in theory)
 [No bing-bing-bing of the molecules colliding with the container!]

- Pressure versus Volume. As $P \uparrow$ Volume \downarrow . The shape of the graph is inverse and $P \propto \frac{1}{V}$ (Boyle's Law)
- Volume also goes to zero at absolute zero!! As the temperature of a gas drops it contracts [put a negative sign on the coefficient of expansion] $1/273$ of its original volume for each degree C or K drop in temperature.
- We now use the triple point of water [think phase diagrams] as the standard by which to calibrate the best gas thermometers.
- Electrical resistance thermometers (actually a *thermistor* which is equivalent to our temperature probe) are still superior.



THERMAL EXPANSION OF SOLIDS AND LIQUIDS

The general concept of matter expanding with increased temperature [except water!] is about to be quantified. We'll begin linear (one dimension or 1-D) and travel to area (2-D) and finish with volume (3-D). Buildings, highways and especially bridges must be constructed with expansion/contraction in mind. The expansion is due to an increased separation between molecules as they gain more energy from an increase in temperature.

- To calculate **linear expansion**: “heat ‘em up and move ‘em out!”

$$\Delta \ell = \alpha \ell_0 \Delta T$$

- $\Delta \ell$ = change in length
- α = coefficient of linear expansion with units of $^{\circ}\text{C}^{-1}$ [per Celsius degree, in English].
USE A POSITIVE SIGN FOR EXPANSION; AND A NEGATIVE SIGN FOR CONTRACTION
- ℓ_0 = original length--that's what a subscript of “o” usually means—“*original*”
- ΔT = change in Kelvin temperature--when the T is capital it's usually “absolute”
- There is a table of coefficients in your text or on the web--no need to memorize any!

- Surface area and volume also change with temperature:

$$\Delta A = \beta A_o \Delta T \quad \text{and} \quad \Delta V = \gamma V_o \Delta T$$

$$\beta = 2 \alpha \text{ (area is 2-D)} \quad \text{and} \quad \gamma = 3 \alpha \text{ (volume is expanding 3-D)}$$

- **Bimetallic strips** are constructed of 2 *different* metals. They have different coefficients of expansions \therefore one expands more (or faster) than the other causing a bend in the strip. This often acts like a switch in thermostats. As the strip cools, it straightens and reconnects the circuit. Brass expands more than steel; these 2 metals are common in bimetallic strips.

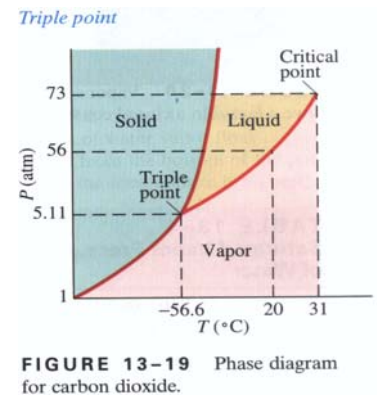
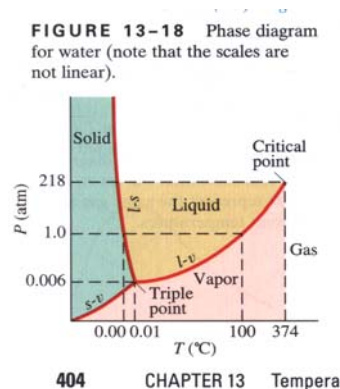


WEIRD WATER

Notice the slope of water's phase diagram versus CO₂'s diagram. All lines represent equilibrium conditions between phases. Triple Point is the T & P at which ALL THREE phases exist in EQUILIBRIUM.

Water achieves its greatest density at 4 °C \therefore ice floats! Antimony & Bismuth also do this.

- This property is weird since the solid phase is not the most tightly packed--for water, that's due to the intermolecular forces between Mickey's ears and chin [remember water looks like Mickey Mouse with the ears being the hydrogens and the lone electron pairs placed either side of Mickey's chin].
- **H-bonding**--the attraction between a H on one molecule and a highly electronegative element (N, O, P, S and of course, F) on an adjacent molecule. DO NOT confuse this with a bonded hydrogen; this is an *intermolecular* force (between molecules).
- Creates a hexagonal solid [ice] crystal structure with a hole in the middle--that's why it's less dense!
- Note the slope of the lines in these phase diagrams. Water slopes to the left MOST substances follow carbon dioxide's pattern.



MACROSCOPIC DESCRIPTION OF AN IDEAL GAS

Recall the gas law unit of study from Chemistry:

Ideal Gas Law:

$$PV = nRT$$

- $n = \# \text{ of moles} = \text{mass (g)}/\text{molar mass} = 6.02 \times 10^{23}$ atoms, molecules, or formula units
 $R = \text{ideal gas constant} = 0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} = 8.31 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ [the “energy R ”]
- $P_1V_1T_2 = P_2V_2T_1$ is the Combined Gas Law
- **Avogadro’s Principle**--Equal volumes of gas @ same T & P contain the same number of molecules
- 1 mole of a gas at STP = 22.4 L = 6.02×10^{23} molecules

PV diagram

Examine the PV diagram.

Notice the plateau for the phase change! The dotted lines represent deviation of a real gas from ideal behavior.

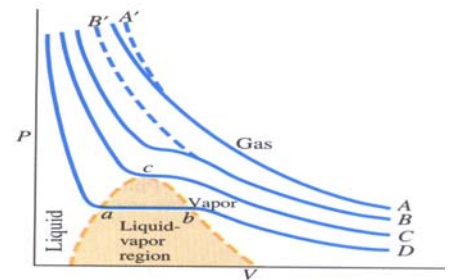


FIGURE 13-17 PV diagram for a real substance. Curves A, B, C, and D represent the same gas at different temperatures.

NEW STUFF (not from chemistry)

- Another way to think about $PV = nRT$ is

$$PV = \left(\frac{N}{6.02 \times 10^{23}} \right) RT$$

Where $N = \text{total \# of molecules}$, so PV can also be expressed in different terms as shown below:

$$PV = N k_B T$$

And, $k_B = \text{Boltzman's constant and MUST} = \frac{\text{"energy"} R}{6.02 \times 10^{23}} = 1.38 \times 10^{-23} \text{ J/K}$

- Don't freak--we'll play a lot of “constants” games--Boltzman's constant is a molecular distribution constant. It just incorporates two other constants, the energy R and Avogadro's number.

KINETIC THEORY OF GASES

Assumptions: (You can also consider these *properties* of ideal gases.)

- Molecules are in constant random motion and travel in a straight line until colliding with each other or the walls of their container
- Molecules have negligible volume; the space between them is what is important
- All collisions are elastic; there is no transfer of momentum
- No attractive forces exist between molecules
- In a pure substance all the molecules are alike
- There are 3 types of motion: vibrational, rotational and translational (from point A to point B). We deal only with translational in the kinetic theory and neglect vibrational and rotational motion.
- What you already know:
 - the more molecules present, the more collisions occurring, the more pressure observed
 - the faster [Temperature = $KE_{avg} = \frac{1}{2} mv^2$] the molecules move, the more likely they are to collide and they collide with more “umph”. (Very scientific word, “umph”).
- New stuff:
 - **Pressure**—equals force per unit area

$$P = \frac{F}{A}$$

$$\text{Temperature} = K_{avg} = \frac{3}{2} k_B T = \frac{1}{2} mv^2$$

You already knew that K_{avg} = temperature and KE; you just didn't know the Boltzman part! There are derivations in your text if you so desire. The 3/2 part simply comes from $\frac{1}{2} mv^2$ having an equal probability of happening on the x , y or z -axis [direction].

HOW FAST IS A GAS MOLECULE? That depends on the temperature!

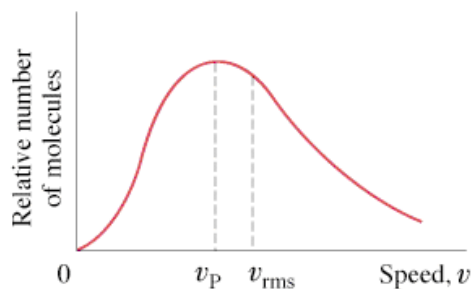
- v_{rms} --is the root mean square velocity also known as rms speed

$$v_{rms} = \sqrt{v^2} = \sqrt{\frac{3k_B T}{m}} = \sqrt{\frac{3RT}{M}}$$

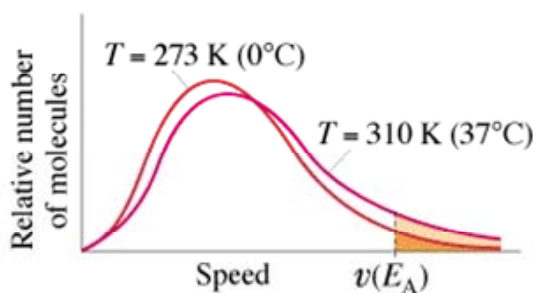
m = molar mass (g/mol) and M = molar mass (kg/mol)

∴ Light, stealth molecules move fast while Heavier, sluggish molecules move slowly.

MAXWELL DISTRIBUTION OF SPEEDS



First, notice that's a funny looking bell curve (normal distribution curve). Why? It's not symmetrical. Note that the v_{rms} is faster than the “most probable speed” at the peak of the curve labeled v_p . You haven't seen the last of these skewed curves.



Recall from biology or chemical kinetics that a 10°C increase in temperature generally doubles the rate of the reaction since more molecules have enough energy to overcome the activation energy, E_A .

So, how does the distribution change when we heat things up? More molecules have sufficient energy to overcome the activation energy, thus the reaction speeds up.

A concept we will use quite a lot is “area under the curve”. Note the shaded regions to the right of the velocity that is associated with E_A . At the warmer temperature, more molecules have sufficient energy to overcome the activation energy, thus the reaction speeds up. As always, “Heat ‘em up. Speed ‘em up.”

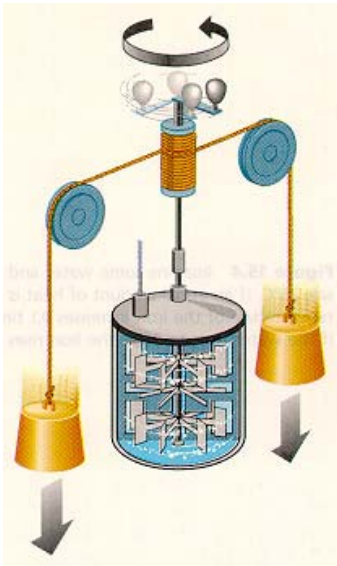
Heat

Thermal equilibrium is established between two objects in contact with each other. Experiments performed by James Joule (1818-1889) showed that whenever heat is gained or lost by a system during some process, the gain or loss can be accounted for by an equivalent quantity of mechanical work done on the system. In English, heat can be transferred from hot to cold and that transfer can perform work.

THE MECHANICAL EQUIVALENT OF HEAT

You already know the value for this concept--you just need all the theory stuff!

- **Heat (or thermal) Energy**--energy that is transferred between a system and its environment because of a temperature difference between them.
 - Units were developed while there was still a big misunderstanding regarding heat! No one understood it was energy and should be measured in energy units! Some of those old units you should be familiar with:
 - **calorie**--the heat required to raise the temperature of 1 gram of water 1 °C
 - **kilocalorie**--the heat required to raise the temperature of 1 kilogram of water 1 °C
 - **Calorie**--really a kilocalorie—it's simply the food industry's way of not depressing us too much when we eat a Snickers bar!
 - **British thermal unit (BTU)**--the heat required to raise the temperature of 1 lb of water from 63 °F to 64 °F. Still used regarding gas grills, ovens, air conditioners, etc.



- **Joule's Experiment**--Joule set out to measure the concept of heat lost due to friction. You probably already know that in mechanical systems, heat is lost and that heat is mostly due to friction. Joule set up the apparatus shown right--a quite simple and elegant one!
- He wanted to measure the mechanical equivalent of heat. The falling masses rotate the paddles, causing the temperature of the water to rise.
- Note the 2 weights are at the same height. If they fall a distance h then they do work on the water when their Potential Energy (mgh) is transferred to the water due to the friction between the water and the paddles as they spin in the water. Since there are 2 weights, \therefore in terms of energy, $2 mgh \propto \Delta T$ [remember that \propto sign means *proportional*--can you feel a constant coming on?] The proportionality constant is the specific heat of water that you already know and love!
- The **mechanical equivalent of heat** is basically a term that refers to the historical importance of Joule's work--without it, we may have never realized that heat is simply energy and thus follows the first law of thermodynamics [law of conservation of energy].

(Told you, you knew it!--Now you just know where it came from!)

$1 \text{ calorie} = 4.186 \text{ J}$

METHODS OF HEAT TRANSFER

- Three methods: ***conduction, convection and radiation.*** We will consider each one separately. NO TRANSFER occurs without a difference in temperature between the 2 objects!
- **conduction**--process of heat transfer involving atomic collisions and contact between 2 objects. Metals are the best conductors since they have free moving electrons in addition to atoms. Ever grabbed a metal pot handle during the cooking process? You usually only do it once! The atoms near the flame are having their vibrational amplitudes increased and they bump into their neighbors increasing their vibrational amplitudes, and so on and so on! Metals do this best since each atom is associated with one or more free electron nearby that also get into the act! Insulators are slow to catch on to the vibrational amplitude game.
- **convection**--process of heat transfer involving differing densities of layers of matter [usually air or water]. Ever heard heat rises? It doesn't. It only moves from hot to cold! What happens say at a campfire, is that the air next to the flame is heated, expands, becomes less dense and floats up to your outstretched, cold hands that are ABOVE the flame. Ever see the road look wet off in the distance in August? Same thing. The air next to the hot road surface is heated, expands and floats. The squiggles you can sometimes see are due to convection currents.
 - Never place hot objects on the balance! It can hurt the balance PLUS you get incorrect readings. Too heavy or too light? Think it through!
- **radiation**--process of heat transfer that ***does NOT involve matter*** but rather, electromagnetic radiation. The other two methods have so far! Back to the campfire--if your hands are above the fire, it's convection. Next to [to the sides of] the fire its radiation.
 - All objects continuously radiate electromagnetic radiation. Infrared radiation is heat transfer by electromagnetic radiation. Dark objects draw radiant heat while light objects reflect radiant heat.

The Laws of Thermodynamics

HEAT AND INTERNAL ENERGY

Thermodynamics is the study of processes in which energy is transferred as heat and work.

- **Internal energy**--symbolized by U ; the sum of ALL energy in a *stationary* piece of matter. Heat, nuclear, chemical, and strain (stretched or compressed springs) energy.
 $U = KE + PE$ (molecular level)
- **Thermal energy**--symbolized by Q ; the *portion* of internal energy that changes when the temperature of the system changes
- **Heat transfer**--caused by temperature difference between system and surroundings
- We will refer to work done *on* or *by* a system ; heat is defined as a transfer of energy due to a temperature difference.
 - **work**--symbolized by W ; is a force acting through a distance; work is also a transfer of energy that is NOT due to a temperature difference
 - Gases cool upon expansion [work done by system] and warm upon compression [work done on system].
 - **open system**--mass and/or energy may enter or leave
 - **closed system**--no mass enters or leaves BUT energy may be exchanged
 - **isolated system**--no energy *in any form* passes across its boundaries

WORK AND HEAT

P, V, T and internal energy establish the *state* of a system.

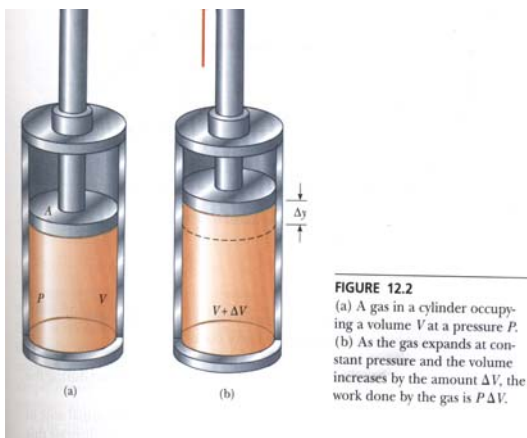


FIGURE 12.2
(a) A gas in a cylinder occupying a volume V at a pressure P .
(b) As the gas expands at constant pressure and the volume increases by the amount ΔV , the work done by the gas is $P\Delta V$.

In equilibrium, the gas contains a volume V and exerts a pressure, P , on the cylinder walls and piston. If the piston has a cross-sectional area, A , the force exerted by the gas on the piston is $F = PA$. Assume the gas expands slowly enough to allow the system to maintain thermodynamic equilibrium at all times. As the piston moves Δy , the work done on the piston by the gas is

$$W = F\Delta y = P(A\Delta y) = P\Delta V$$

[remember the volume of a cylinder is (area of the circle) \times (height)]

$$\text{Work} = W = -P\Delta V$$

About those pesky sign conventions...

In AP Physics B, W is defined as “work done ON a system” and is (+). Well, that makes sense since work is literally “added” to the system. So, it stands to reason that when work is done BY a system (like an expanding gas moving the walls of its container), then W should bear a (-) sign. In the case of an expanding gas the “work” term is simply $P\Delta V$, so it bears a (-) sign.

The sign of Q is just like the sign of ΔH in chemistry. Exothermic, heat exited the system; this is detected by an increase in temperature and represented by $-Q$ or $-\Delta H$. The converse is also true. If heat was added to the system (from the surroundings), it is detected by a decrease in temperature, is referred to as endothermic and represented by $+Q$ or $+\Delta H$.

In summary, the scoop on the signs is that if “it” was “added” to the system, “it” bears a (+) sign. If “it” was removed or exited the system, then “it” bears a (-) sign.

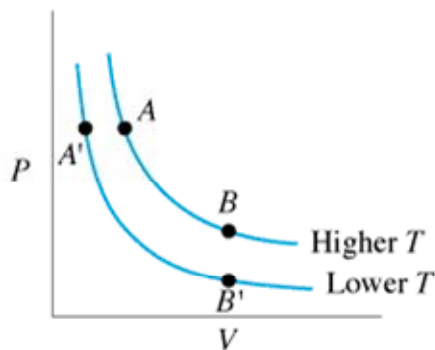
- **expansion**--work is done *by* the gas (system) and the gas cools; W is negative since energy is leaving the system [since the gas cools, Q is also negative]
- **compression**--work is done *on* the system and the gas warms; W is positive since energy is increasing [Q is also positive since the gas is warmed, more KE added to the system]
- **Heat added is + Heat lost is (-) Work ON system is (+) Work BY system is(-)**
Think of it as “adding” either heat or work TO the system being studied. [Either increases its energy.]

FIRST LAW OF THERMODYNAMICS APPLIED TO SOME SIMPLE SYSTEMS

The law of conservation of energy IS the first law of thermodynamics. Mathematically stated:

$$\Delta U = Q + W = \frac{3}{2}nRT$$

- **isothermal process**-- $\Delta T = 0$ since T is constant; process carried out at constant temperature and the curves are called *isotherms*. Consider the following graph: isotherms are for A, B and A', B'



$PV = nRT$ or better yet, $P = (nRT)/V$ in $y = mx + b$ format, so if T is constant then $P = (\text{all constants combined})/V$.

As P increases, V decreases so the relationship is inverse, BUT there is a constant inverse relationship. You just get a new constant at each temperature!

- **heat reservoir**--a body whose mass is so large that, ideally,

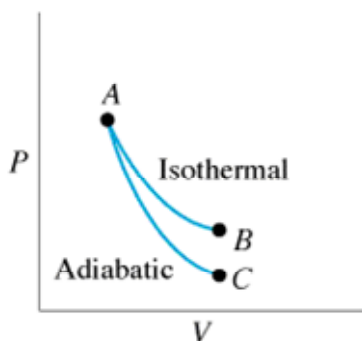
its temperature does not change significantly when heat is exchanged with a system (think oceans!)

- Think back to the piston system. Assume heat is added and the gas expands slowly enough to allow the system to maintain thermodynamic equilibrium at all times. As the piston moves Δy , the work done on the piston by the gas is $-P\Delta V$
 - Now apply $\Delta U = Q + W$
 - T is constant, $\Delta T = 0$; AND $\frac{3}{2} nRT = \Delta U = 0$ since internal energy does NOT change since there is NO heat exchange! That means that

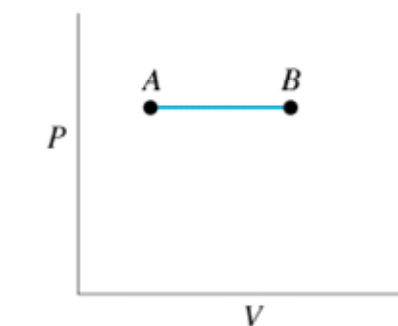
$$0 = Q + W \quad \text{So...}$$

$$Q = -W \quad \text{when } T \text{ is held constant and the work done BY the gas in an isothermal process leaves the system [hence the negative sign] and equals the heat added to the gas [in magnitude]}$$
- **adiabatic process**--No heat is allowed to flow into or out of the system $\therefore Q = 0$
 - either this change happens very quickly and the heat can't flow fast enough OR
 - the system is extremely well insulated--the rapid expansion of gases in the internal combustion engine is a good example--it's nearly adiabatic
 - If $Q = 0$ AND $\Delta U = Q + W$, then $\Delta U = +W$
 - The internal energy decreases with an expansion and increases with a compression as does Temperature
- **isobaric process**-- "same pressure"; $P = \text{constant}$, $W = -P\Delta V$
- **isochoric [or isovolumetric] process**-- "same volume"; $V = \text{constant}$, $W = 0$
- **isolated system**-- $Q = |W| = 0$, $\Delta U = 0$ and so $U_{\text{initial}} = U_{\text{final}}$; internal energy remains constant!
- It is also important to point out that Q & W depend on the path taken in the process, BUT the quantity $Q - W$ is *independent* of the path taken.

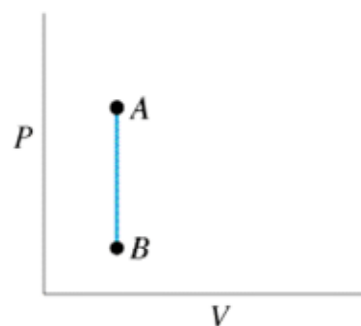
YOU MUST KNOW THESE GRAPHS AND WHICH GOES WITH WHICH PROCESS!!



for Adiabatic (curve AC)
 $Q = 0$ and $\Delta U = Q + W$
 $\therefore \Delta U = W$



for Isobaric $P = \text{constant}$
 $\therefore W = -P\Delta V = -(\text{area under the curve})$
 $\therefore \Delta U = Q + (-P\Delta V)$



for Isochoric (Isovolumetric) $\Delta V = 0$
 $\therefore W = -P\Delta V = 0$
 $\therefore \Delta U = Q + 0 = Q$

HEAT ENGINES AND THE SECOND LAW OF THERMODYNAMICS

A heat engine is a device that converts thermal energy to either mechanical or electrical energy. Coal is burned, water turns to steam which causes a turbine to turn which drives an electrical energy. Refrigerators are the opposite of heat engines!

- **heat engine**--carries some working substance through a cyclic process during which
 - heat is absorbed from a source at a higher temperature
 - work is done by the engine
 - heat is expelled by the engine to a reservoir at a lower temperature ["heat sink"]
- Q_h --the engine absorbs a quantity of heat [boiler on Titanic]
- W --work done BY the expanding steam to turn a gizmo [propeller on Titanic]
- Q_c --heat given up by the engine to heat sink [pipe system to return "spent" water]
- T_c --temperature of heat sink
- Since the working substance [water] goes through a cycle its initial and final energies are equal
- $\Delta U = 0$ AND $Q = -W$ from the first law of thermodynamics
- That means net W done BY [leaving] a heat engine is equal to the net heat flowing into it.
- $Q_{\text{net}} = Q_h - Q_c = W$; If the working substance is a gas--the net work done for a cyclic process is the area enclosed by the curve representing the process on a P-V diagram.

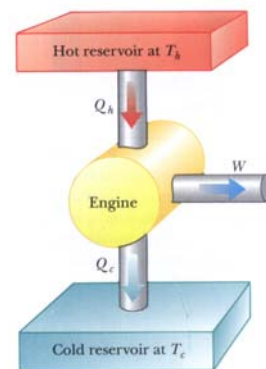


FIGURE 12.6
A schematic representation of a heat engine. The engine receives heat Q_h from the hot reservoir, expels heat Q_c to the cold reservoir, and does work W .

THERMAL EFFICIENCY

Symbolized by e and equal to the ratio of the net work done to the heat absorbed at the higher temperature during one cycle.

$$\text{efficiency, } e = \frac{|W|}{Q_H} = \frac{Q_H - Q_C}{Q_H} = 1 - \frac{Q_C}{Q_H}$$

Think of it as the ratio between what you get [mechanical energy] and what you give [thermal energy].

- 100% efficiency is achieved ONLY if $Q_c = 0$. That means that no heat was lost in the conversion--IMPOSSIBLE!!
- The first law states: Energy cannot be created nor destroyed but only change forms [you cannot get more out than you put in]. The second law shows us that E changes are not 100% efficient [you cannot break even]. That's why you've learned that Energy is lost usually as heat. The heat engine studies let you defend it and one more idea.....
- Perpetual motion machines--one that allows for 100% efficiency OR allows you to get more Energy out than you put in. NOT GOING TO HAPPEN! There is a patent issued on one in America, though.

REVERSIBLE AND IRREVERSIBLE PROCESSES

Reversible is theory. Frictionless, no turbulence, etc.
Irreversible is reality in the natural world.

THE CARNOT ENGINE

A theoretical engine that operates in an ideal, reversible cycle--called a Carnot cycle--between two reservoirs is the most efficient engine possible.

- **Carnot's theorem**--No real engine operating between two heat reservoirs can be more efficient than a Carnot engine, operating between the same two reservoirs.
- The thermal efficiency of a Carnot engine is:

$$e_c = \frac{T_h - T_c}{T_h} = 1 - \frac{T_c}{T_h}$$

All of the T 's are in Kelvin and the only way e can equal 100% is if $T_c = 0$ K.

I've always heard that the reason scientists have never reached absolute zero is due to the lack of a heat sink--you've got to have a place colder than absolute zero to put the heat on that last transition from 1 K to 0 K!! Since all molecular motion stops at absolute zero, you'd have to have a substance/place with negative molecular motion--antimatter??? A problem for far greater minds than mine!

ENTROPY

The amount of disorder in a system. Calculated as follows:

$$S = \frac{\Delta Q_r}{T}$$

SUMMARY:

There are many ways to state the laws of thermodynamics. Here are a few!

Zeroth law: If two systems are in thermal equilibrium with a third system, then they are in thermal equilibrium with each other.

First law: $\Delta U = Q - W$ or the statement of the law of conservation of energy.

Second law: Heat flows spontaneously from a hot object to a cold one, but not the reverse.
No heat engine operating in a cycle can absorb thermal energy from a reservoir and just perform an equal amount of work.
There cannot be a 100% efficient heat engine--one that can transform all heat into work.
The entropy of the universe is always increasing

Third law: The entropy of a pure perfect crystal is zero at zero degrees Kelvin.