

Review for Final Exam: Giancoli Chapters 17, 19, 20

Physics 1422 (Dr. Yost)

The final exam will be comprehensive, including all material on the previous exams, plus chapters 17, 19, and 20 of the Giancoli text. This review just includes the last three chapters. See the earlier reviews as well. You may use any calculator for this exam. You will be given any constants you need, and page of notes (any size, front and back). Equations will not be provided for the final exam, but constants will.

Chapter 17: Temperature, Thermal Expansion, and the Ideal Gas Law

Concepts: microscopic states, macroscopic states, state variables, temperature, thermometers, Celcius, Fahrenheit, thermal equilibrium, zeroth law of thermodynamics, thermal expansion, equation of state, absolute zero, Kelvin scale, Ideal Gas Law, atoms, molecules, atomic or molecular mass, Avogadro's number, moles

Equations:

Two systems (collections of atoms or molecules) are in thermal equilibrium if they are in thermal contact, so that heat could flow between them, but no heat flows.

Zeroth law of thermodynamics: two systems in thermal equilibrium with a third system are in thermal equilibrium with each other.

Temperature is defined so that two systems in thermal equilibrium will have the same temperature. It can be defined practically by measuring the changes in materials which depend on temperature.

Linear expansion: $\Delta L = \alpha L_0 \Delta T$.

Volume expansion: $\Delta V = \beta V_0 \Delta T$.

The factors α and β are the coefficients of linear and bulk expansion, respectively. Normally, $\beta \approx 3\alpha$. These coefficients have dimensions of $(C^\circ)^{-1}$.

The ideal gas law is $PV = nRT = NkT$, where P and V are the pressure and volume of the gas, T is the absolute temperature, n is the number of moles, N is the number of atoms or molecules, and the constants R and k are universal:

Boltzmann constant: $k = R/N_A = 1.38 \times 10^{-23}$ J/K. (This is effectively a unit conversion factor between Kelvins and Joules.)

Ideal gas constant: $R = N_A k = 8.315 \text{ J}/(\text{mol} \cdot \text{K})$, where N_A is Avogadro's number (see Units below). The number of moles is $n = N/N_A$.

The ideal gas law applies to any gas which is sufficiently dilute that the atoms or molecules can be considered as non-interacting.

Units:

Temperature is usually measured in Celsius and Kelvin scales, with T (Kelvin) = T (Celsius) + 273.15K.

Atomic mass units (u) are used to measure the masses of single atoms and molecules. This is defined so that an atom of ^{12}C has an atomic mass of exactly 12 u.

1 gram = N_A u, where $N_A = 6.02 \times 10^{23}$ is called Avogadro's number.

A mole is defined to be N_A particles, *i.e.* the number of particles of atomic weight M u which it takes to make up M grams of the substance. 1 u = 1 g/mol.

Chapter 19: Heat and the First Law of Thermodynamics

Concepts: Heat, energy transfer, calories, thermal energy, internal energy, specific heat, latent heat, first law of thermodynamics, closed system, open system, isolated system, isothermal, quasistatic, isochoric, isobaric, adiabatic, molar specific heats of gases, equipartition of energy, heat transfer, conduction, convection, radiation, solar constant

Sections skipped: 4 (Calorimetry – Solving Problems)

Equations:

The internal energy of an ideal gas is $U = \frac{d}{2}nRT$, where d is the number of degrees of freedom of the gas. For a monatomic gas, $d = 3$ always. For a diatomic gas, rotations are normally possible at room temperature, raising d to 5, but at very low temperatures, the rotational degrees of freedom can freeze out, leaving $d = 3$ again. At very high temperatures, vibrational modes may raise d to 7. However, the vibrational modes only become active when it is almost hot enough to break apart the molecule, so these modes are not seen for a wide range of temperatures.

The amount of heat Q needed to change the temperature of a mass m of a material by ΔT is given by $Q = mc\Delta T$, where c is the specific heat.

The specific heat of water is 1 cal/g°C.

The amount of heat Q needed to evaporate a mass m of liquid is $Q = mL_v$ where L_v is the latent heat of vaporization. The amount of heat Q which must be removed from a mass m of liquid to freeze it is $Q = mL_f$, where L_f is the latent heat of fusion.

The molar specific heats are defined by $C = Mc$, where M is the molecular mass (grams per mole). The molar specific heat of an ideal gas at constant volume is $C_v = \frac{d}{2}R$, while at constant pressure it is $C_p = (\frac{d}{2} + 1)R$, where d is the number

of degrees of freedom ($d = 3$ for monatomic gases). The internal energy of a gas is $U = nC_vT$. The internal energy always depends only on the temperature. Note that it is always proportional to C_v , regardless of the process. The first law of thermodynamics states that for a closed system (no material enters or leaves, but energy may), the change in internal energy is

$$\Delta U = Q - W. \quad (1)$$

where Q is the heat flow into the system, and W is the work done *by* the system.

The work done by a gas in a closed system is $W = \int PdV$. This is the area under the curve on a PV diagram. The net work done in a closed cycle on a PV diagram is the area inside the curve traversed by the system through the cycle.

For ideal gasses:

Process	Definition	ΔU	W	Q
Isochoric	$\Delta V = 0$	$nC_v\Delta T$	0	$nC_v\Delta T$
Isobaric	$\Delta P = 0$	$nC_v\Delta T$	$P\Delta V = nR\Delta T$	$nC_p\Delta T$
Isothermal	$\Delta T = 0$	0	$nRT \ln(V_2/V_1)$	W
Adiabatic	$Q = 0$	$nC_v\Delta T$	$-nC_v\Delta T$	0

- Isothermal process: $PV = \text{constant}$.
- Adiabatic process: $PV^\gamma = \text{constant}$ where $\gamma = C_p/C_v$.

If a system is taken through a series of PV transformations that make a closed cycle, ending at the initial conditions, then $\Delta U = 0$ for the complete cycle, and the net heat input equals the net work done by the system. $Q = W$ is the area enclosed by the curve on a PV diagram.

The heat conducted per unit time across a distance l separating two temperatures T_1 and T_2 , through a cross-sectional area A is

$$\frac{dQ}{dt} = \frac{kA}{l}(T_1 - T_2) \quad (2)$$

where k is the thermal conductivity.

Heat can be transferred by electromagnetic radiation. An object with surface area A at absolute temperature T radiates energy

$$\frac{dQ}{dt} = e\sigma AT^4 \quad (3)$$

where e is a number between 0 and 1 called the “emissivity”, and $\sigma = 5.67 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4$ is the Stefan-Boltzmann constant. Objects with high emissivity are

the best absorbers and also the best emitters. Such objects are black, and perfect absorbers ($e = 1$) are called blackbodies. The radiation at room temperature is mostly in the infrared wavelengths. Shiny objects are the worst absorbers since they reflect most of the radiation striking them. Therefore, they are also the worst emitters. A perfect reflector would have emissivity 0.

If the body at temperature T_1 is surrounded by an environment at temperature T_2 , then the energy radiated into the environment is

$$\frac{dQ}{dt} = e\sigma A(T_1^4 - T_2^4). \quad (4)$$

The Earth receives a power per unit area of 1350 W/m^2 from the sun. This is called the solar constant. When the sun is at an angle θ in the sky with respect to the zenith (overhead), the power received is $1350 \text{ W/m}^2 \cos \theta$. About 74% of this power actually reaches the surface of the earth (1000 W/m^2), and a fraction e will be absorbed by an object with emissivity e .

Units:

1 calorie (cal) = 4.186 J is the amount of heat needed to raise the temperature of 1 g of water by 1°C . 1 kcal is also called a “large calorie” (Cal). Food calories are kilocalories (the amount of energy which may be obtained by breaking the chemical bonds in the food).

Chapter 20: Second Law of Thermodynamics

Concepts: heat engines, direction of heat flow, efficiency, second law of thermodynamics, reversible and irreversible processes, Carnot cycle, refrigeration, entropy, order and disorder, availability of energy, statistical nature of entropy, third law of thermodynamics, absolute zero

Sections skipped: There are no quantitative problems from 7 – 10, but these sections contain useful conceptual information for understanding entropy.

Equations:

Heat engines operate between a high temperature T_H and low temperature T_L , taking in heat Q_H at the high temperature, exhausting heat Q_L at the low temperature, while doing work $W = Q_H - Q_L$. The performance of an engine is measured by its efficiency: the work done per unit of heat put into the engine, $e = W/Q_H$.

Refrigerators (or air conditioners) are heat engines operated in reverse, heat Q_H is exhausted at the higher temperature, while heat Q_L is taken in at a cooler temperature. Work $W = Q_H - Q_L$ must be done on the refrigerator to accomplish this. Since refrigerators are designed to transfer as much heat (Q_L) out of a cold place as possible with a given amount of work, the efficiency is measured by the coefficient of performance defined as $\text{CP} = Q_L/W$.

Heat pumps are refrigerators whose purpose is to transfer as much energy as possible

into a warmer place for a given amount of work. Therefore, their coefficient of performance is defined as $CP = Q_H/W$ instead.

In each case, the ideal engine, refrigerator or heat pump is constructed from the Carnot cycle, which maximizes the efficiency of coefficient of performance. For the Carnot cycle, $Q_H/Q_L = T_H/T_L$ in terms of absolute temperatures. Substituting this into the expressions for e or CP gives the maximum efficiency or coefficient of performance allowed by the second law of thermodynamics.

In summary, the measures of efficiency are:

System	Heat Engine	Refrigerator	Heat Pump
Measure of Performance	$e = \frac{W}{Q_H}$	$CP = \frac{Q_L}{W}$	$CP = \frac{Q_H}{W}$
Carnot (Ideal) Case	$e_C = \frac{T_H - T_L}{T_H}$	$CP_C = \frac{T_L}{T_H - T_L}$	$CP_C = \frac{T_H}{T_H - T_L}$

If an tiny quantity of heat dQ flows into a system at temperature T , the entropy S of the system changes by dQ/T . For any *reversible* process between states a and b , the change in entropy is

$$\Delta S_{ab} = \int_a^b \frac{dQ}{T}. \quad (5)$$

This does not depend on the process, so the entropy is purely a property of the states, *i.e.* a state variable. In other words, entropy depends on what point a system is at on a PV diagram, but not on how it got to that point. To calculate the change of entropy during an irreversible process, you may apply the previous equation to *any reversible process* which connects the same initial and final state. In any natural process, $\Delta S > 0$ when the entire environment is included. Adiabatic processes are also isentropic, $\Delta S = 0$.

Process	ΔS
Isochoric	$nC_v \ln \left(\frac{T_2}{T_1} \right)$
Isobaric	$nC_p \ln \left(\frac{T_2}{T_1} \right)$
Isothermal	$nR \ln \left(\frac{V_2}{V_1} \right)$
Adiabatic	0

For an ideal gas: $S = nR \ln(V/V_0) + nC_V \ln(T/T_0) + S_0$ where S_0 is the entropy at a reference state T_0, V_0 . This is normally used to compare the entropy for two different states, so the value of S_0 cancels out and does not need to be specified.

Entropy can be defined statistically in terms of the number of microscopic states, or *microstates*, corresponding to a given set of thermodynamic variables, or *macrostate*. If this number is \mathcal{W} , the entropy is $S = k \ln \mathcal{W}$.

In practice, calculating \mathcal{W} can be difficult, but sometimes comparisons are possible. For example, increasing the volume of a container by a factor of r increases the number of places any atom can be by a factor of r , and the number of possible positions for all N of the atoms increases by a factor of r^N . Then the entropy increases by $\Delta S = k \ln(r^N) = Nk \ln r = nR \ln r$. If $r = V_2/V_1$ is the ratio of two volumes, and all else is held fixed, $\Delta S = nR \ln(V_2/V_1)$. This is consistent with the above equation for the entropy of an ideal gas.

The second law of thermodynamics can be stated in a variety of equivalent ways:

1. Heat flows spontaneously from a hot object to a cold one, but not in reverse.
2. No device can transform a given amount of heat completely into work.
3. No device is possible whose sole effect is to transfer heat from a system at a lower temperature to a system at a higher temperature.
4. The entropy of an isolated system never decreases.
5. Natural processes move toward a state of greater disorder, when the entire environment is taken into account.

The third law of thermodynamics states that absolute zero cannot be reached by any finite number of thermodynamic processes. A consequence of the second and third law together is that there can be no 100% efficient heat engine or refrigerator. An alternative statement of the third law is that $S = 0$ at $T = 0$.