Review for Final Exam: Giancoli Chapters 19 – 20

Physics 1422 (Dr. Yost)

The final exam will be comprehensive, including all material on the previous exams, plus chapters 19 and 20 of the Giancoli text. This review just includes the last two chapters. See the earlier reviews as well. You may use any calculator for this exam. You will be given any constants you need, and you may bring one page of handwritten notes (any size, front and back, no photocopies). Equations will not be provided for the final exam, but constants will. You should put any equations you feel you need in your notes.

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^{o}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

$$\Delta U = Q - W. \tag{1}$$

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: $PV =$ constant.
- Adiabatic process: $PV^\gamma =$ 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

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

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

$$\frac{dQ}{dt} = e\sigma AT^4 \tag{3}$$

where e is a number between 0 and 1 called the "emissivity", and $\sigma = 5.67 \times 10^{-8}$ W/ m$^2\cdot$ 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

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

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 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^oC. 1 kcal is also called a "large calory" (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: 7 – 10

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 CP $= Q_L/W$.

Heat pumps are refrigerators whose purpose is to tranfer 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	$e = \frac{W}{Q_H}$	$\mathrm{CP} = \frac{Q_L}{W}$	$\mathrm{CP} = \frac{Q_H}{W}$
Carnot (Ideal) Case	$e_\mathrm{C} = \frac{T_H - T_L}{T_H}$	$\mathrm{CP_C} = \frac{T_L}{T_H - T_L}$	$\mathrm{CP_C} = \frac{T_H}{T_H - T_L}$

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

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

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

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, $\Delta S = 0$.

For an ideal gas: $S = nR\ln(V/V0) + nC_V \ln(T/T0) + S_0$ where S_0 is the entropy at a reference state T_0, V_0.

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.