

Exam I is on **Tuesday the 20th of September, 7:55-8:55 am** in the usual meeting room. It is closed book and over Chapter 2 and a small part of Chapter 3. ***You may have one 4 x 6" sheet (both sides) of handwritten information of your choosing for use on the exam.***

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## **Objectives**

### **Know or be able to:**

#### Chapter 2.

1. Definitions of system, surroundings, heat and work.
2. State the First Law of Thermodynamics in words and with an equation involving heat and work, defining all terms used precisely. Know how changing the definitions changes the equation.
3. Meaning of and distinction between *state* variables and *path* variables.
4. The ideal gas equation of state. What is an ideal gas, and why its energy does not depend on volume.
5. General expression for expansion work. Calculate  $w$ , given  $p_{\text{ex}}$  and volume change for constant external pressure.
6. Convert  $p$  between atm, bars and Pa; convert  $pV$  in L bar (or L atm) to joules and calories.
7. That energy and enthalpy of ***any ideal*** gas depend only on temperature by the relations:  $\Delta U = C_v \Delta T$  and  $\Delta H = C_p \Delta T$  for ***any*** process of an ideal gas, regardless of whether  $V$  and  $p$  are constant.
8. Definition of enthalpy and the reason for its use. What 3 things must be true so that  $\Delta H = q$ ?
9. Definitions of  $C_v$  and  $C_p$ . How to use these to calculate energy and enthalpy changes during heating and cooling.
10. Relation between  $C_v$  and  $C_p$  for an ideal gas, and why they differ.
11. Difference between a reversible and irreversible expansion and compression. How the magnitude of work is related to the reversibility.

(over)

12. Calculate  $\Delta U$ ,  $\Delta H$ ,  $q$ , and  $w$  for isothermal expansions of ideal gas for: (1) reversible, (2) constant pressure, and (3) free expansions (external pressure = 0). Also for adiabatic expansions when adequate information is given.
13. Definition of **partial pressure** and **vapor pressure**; relationship between boiling point of a liquid and its vapor pressure; precisely what is boiling?
14. Definition of standard states. (Note that 25°C is NOT part of the definition).  
The more recent convention is to use 1 bar instead of 1 atm for the standard state.
15. Definition of heat of formation. (In other words, know what chemical reaction has  $\Delta_r H^\circ$  equal to  $\Delta_f H^\circ$  for a given compound. Use  $\Delta_f H^\circ$  tables to calculate the  $\Delta_r H^\circ$  for a chemical reaction or a phase change.
16. Use the table of bond dissociation energies (D) to *estimate*  $\Delta H$  for a chemical reaction, given the bonding arrangements. Understand why the signs are reversed relative to when using  $\Delta_f H^\circ$ s.
17. The *approximate* magnitudes (within a factor of 10 or so) of energy or enthalpy changes (in kJ/mol) accompanying the following processes: (1) heating 1 kg of something by 100 degrees C. (2) a molar phase change; (3) a typical chemical reaction producing a mole of product. (*Note: this is the chemical equivalent of asking "Is the distance to the SUB about 1 mile, 1 block, or 10 feet from our lecture room?"*)

Chapter 3 (pages 55-60, but primarily Equations 3.5-3.8, as summarized below)

1. Second Law of thermodynamics in words and with an equation.
2. The operational definition of entropy change of a system,  $\Delta S$  (how it is measured or calculated), i.e.,  $dS = dq_{\text{rev}}/T$  for *any* process, and  $\Delta S = q_{\text{rev}}/T$  for an isothermal process.
3. How to determine  $\Delta S_{\text{surr}}$ , the entropy change of the surroundings. ( $\Delta S_{\text{surr}} = -q/T_{\text{surr}}$ ).

### General

Be able to work problems of the type given as homework for Chapter 2.