Preface
Our last in-class midterm, scheduled for Friday draws upon our study of Chapters 11, 12, and 13 in Burdge & Overby, *Chemistry: Atoms First*.

*Not covered:* Ch.11: Molecular speed, Diffusion and effusion (§11.2, pp.416-418); Calculation and measurement of pressure (§11.3, pp.419m-421); Real gases (all of §11.6, pp.434b-438b). Ch.12: Surface tension, viscosity, Clausius-Clapeyron equation (§12.2, pp.470-471b and pp.473m-475t); Crystal structure, Types of crystals, Amorphous solids (all of §12.3, §12.4, and §12.5, pp.475b-488m). Ch.13: Concentration units (all of (§13.3, pp.516-519t); Colligative properties, Colloids (all of §13.5 and §13.6, pp.521b-537m).

*Information provided* (see below): Periodic table, equations (see below); any necessary constants or tabulated data

*Assumed knowledge:* Proper treatment of significant figures in calculations; prefixes used with SI units: kilo- (k) $10^3$, centi- (c) $10^{-2}$, milli- (m) $10^{-3}$, micro (µ) $10^{-6}$, nano- (n) $10^{-9}$, pico- (p) $10^{-12}$. Ability to use the periodic table and general periodic trends (*e.g.* electronegativity) in chemical reasoning and problem-solving. Names and formulas for all elements, diatomic molecules, polyatomic ions, associated acids, and other molecules from the *Mandatory Chemical Vocabulary* handout. Note that you need to know the correct charges for all ions.

*Important!* Be sure you know how to obtain a valid Lewis structure from a chemical formula or given skeletal structure and how to interpret Lewis structures in deducing molecular geometry and polarity.

Be sure to review example problems from class and online problem assignments. Many examples, exercises and other text resources, are listed below with the learning goals for this part of the course, should you wish additional opportunities to practice on specific topics.

**Learning Goals**

Exam questions are intended to assess your attainment of a set of learning goals for each chapter. These are compiled below along with relevant text resources.

**Chapter 11 – Gases**

- Understand the basic properties of gases and how they differ from other phases of matter
  
  Online HW 11, Questions 1, 2

- Understand the postulates and results of kinetic molecular theory
  
  See the discussion on p.415, also Fig.11.4 (a) (p.416)

- Understand what pressure is and how to convert between units of pressure
  
  Section Review 11.3.1 and 11.3.2 (p. 421); online HW 11, Question 4

- Relate pressure, volume, temperature, and number of moles for an ideal gas
  
  This includes applications of the simple gas laws, combined gas law, and ideal gas equation. See Worked
• Use and understand Dalton’s law of partial pressures
  See Worked Examples and appended Practice Problems 11.12 and 11.13, and Section Review for §11.7 (p.442); online HW 11, Questions 12-15, 19, 20, 21. See also Key Skills: Mole fractions (pp.462-463).

• Relate what you know about gases to stoichiometry and related problems
  See Worked Examples and appended Practice Problems 11.14 - 11.16, and Section Review for §11.8 (p.449t); online HW 11, Questions 17-19, 21.

Chapter 12 – Intermolecular Forces

• Understand the different types of intermolecular forces and when they would occur
  See Worked Example 12.1 and appended Practice Problems, and Section Review for §12.1 (p.469b); online HW 12, Questions 2-5. Note as well how these forces apply to mixtures (see for example, Fig.13.3, p.513, and online HW 13, Question 1.

• Understand the molecular basis of phase changes
  See the discussion on pp.471b-472 and Fig.12.12, pp.490-492; online HW 12, Questions 1, 12.

• Understand the information contained in a heating/cooling curve
  See Worked Example 12.7 and appended Practice Problems, and Section Review for §12.6 (p.494t); online HW 12, Questions 6-8.

• Understand the information contained in a phase diagram
  See Worked Example 12.8 and appended Practice Problems, and Section Review for §12.7 (p.496b), online HW 12, Questions 9-11.

Chapter 13 – Physical Properties of Solutions

• Determine when and why substances are likely to dissolve
  See Worked Example 13.1 and Section Review for §13.2 (p.515b); online HW 13, Questions 2, 3, 5-10.

• Understand the concept of entropy and its role in determining the chemical potential energy of a system. In particular, describe how entropy favors dissolution of solids and disfavors dissolution of gases
  See discussion pp.513b-514 and Key Skills: Entropy as a driving force, pp.548-549; online HW 13, Question 4.

• Define the term miscible (see p.514)

• Understand the effects of temperature and pressure on solubility (§13.4)
  See the discussion on pp.519-521; Worked Example 13.4 and Practice Problems A and B; Online HW 13, Questions 11-13.
Vapor pressure, H\textsubscript{2}O(l)

<table>
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<th>Temperature (°C)</th>
<th>(P_{\text{H}_2\text{O(g)}}) (torr)</th>
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<tr>
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<td>95</td>
<td>633.9</td>
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<tr>
<td>100</td>
<td>760.0</td>
</tr>
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</table>

Equations and information

Physical constants, conversion factors:
Avogadro’s number: \(N_A = 6.022 \times 10^{23}\)
Gas constant, \(R\): 0.08206 L·atm/mol·K
Absolute zero (absolute temperature scale): 0.00 K = 273.15 °C
Pressure: 1 atm = 14.7 psi = 760 torr = 1.01325 bar

Ideal gas equation:
\[PV = nRT\]

Dalton’s law of partial pressures
\[P_{\text{total}} = \Sigma P_i\]

Mole fraction: \(\chi_i = n_i/n_{\text{total}}\)

Partial pressure and mole fraction: \(P_i = \chi_i \cdot P_{\text{total}}\)
Sample exam questions

Many of the following sample questions have appeared on previous exams. The selection of questions appearing below do not necessarily reflect the weight or distribution of topics that will appear on the actual exam.

1. Perform the following conversions:

   (a) 5.0 atm to psi  
   (b) 37.27 °C to K  
   (c) 5.55 mol of an ideal gas at STP to volume in L

2. The volume of a sample of an ideal gas held at a fixed temperature is 2.25 L under at a pressure of 8.46 atm.
   (a) What is the volume of the gas if it expands at constant temperature until the pressure is 1.75 atm?
   (b) If the initial temperature of the gas is 298.15 K, by how much should the temperature be decreased at constant volume in order to achieve the same final pressure as in part (a)?

3. The average kinetic energy of the molecules in an ideal gas depends on
   (a) the molecular mass of the molecules
   (b) the temperature of the gas
   (c) the density of the gas
   (d) all of the above
   (e) none of the above

4. True or false (in each case, check one):

   (a) The pressure of an ideal gas is inversely proportional to temperature (constant \( n, V \)).  True ____    False____

   (b) The partial pressure of a component of an ideal gas mixture is proportional to its mole fraction.  True ____    False____

   (c) Dispersion forces are stronger than dipole-dipole forces.  True ____    False____

   (d) Energy transfer to a pure substance always raises its temperature.  True ____    False____

   (e) The entropy of the liquid phase of a substance is greater than its solid phase.  True ____    False____
5. (a) A balloon contains 0.133 mol of an ideal gas which occupies a volume of 3.00 L. If an additional 0.067 mol of gas is added to the balloon (at the same temperature and pressure), what will its final volume be? (b) If $P = 1$ atm (exactly) what is the temperature of the gas samples described in part (a)?

6. A chemist investigating a gas-evolution reaction wants to determine the identity of the gas produced, believing that the gas must be one of three possibilities: oxygen, nitrogen, or nitrogen dioxide. She collected a 1.00-L sample of the gas at standard temperature and pressure (STP) and measured its density, obtaining 1.25 g/L. What should the chemist conclude about the identity of the gas based on this information?

7. (a) For the reaction between hydrogen gas and nitrogen gas to produce ammonia:

$$\text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g)$$

If 4.00 L of nitrogen is mixed with 8.00 L of hydrogen at the same temperature and pressure, which is the limiting reactant?

(b) If the reaction above takes place at 117 °C with the reactant volumes given and goes to completion, how much ammonia, in mol, is produced if its partial pressure is exactly 10 atm?
8. In order to determine the rate of photosynthesis, the oxygen gas produced by a species of blue-green algae was collected over water at a temperature of 293 K and a total pressure of 755.2 torr. If a total of 1.02 L of gas was collected, what mass (in g) of oxygen gas was formed?

9. Calculate the energy required to convert exactly 2 mol ice at −10.00 °C to steam at 110.00 °C. Use the following data: Molar heat capacities: H₂O(s) 37.7 J/mol·°C; H₂O(l) 75.3 J/mol·°C; H₂O(g) 35.85 J/mol·°C; enthalpy of fusion, H₂O: 6.01 kJ/mol; enthalpy of vaporization, H₂O: 40.79 kJ/mol.

10. For the phase diagram shown at right, label axes, assign phases to regions (1), (2), and (3), and draw a straight-line arrow showing the conversion of liquid to solid at constant temperature. What does the line B-C represent?
11. Consider the pure substances made up of each of the following molecules.

(i) \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--H} \\
\text{H--O--H}
\end{array}
\]

(ii) \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--H} \\
\text{H--O--H}
\end{array}
\]

(iii) \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--H} \\
\text{H--O--H}
\end{array}
\]

(iv) \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--H} \\
\text{H--O--H}
\end{array}
\]

(v) \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--H} \\
\text{H--O--H}
\end{array}
\]

(a) For which pure substances are its molecules able to form hydrogen bonds?

(b) Given the types of intermolecular forces possible in each case, predict the order of boiling points, **lowest to highest**.

12. Which molecule has the higher boiling point? Explain your answer.

a. \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--H} \\
\text{H--C--H}
\end{array}
\]

b. \[
\begin{array}{c}
\text{H} \\
\text{H--C--C--C--C--C--H} \\
\text{H--C--C--H--H}
\end{array}
\]