Preface
The third in-class midterm, scheduled for Friday draws upon our study of Chapters 8, 9, and 10 in Burdge & Overby, Chemistry: Atoms First.

Not covered: Ch.8: Periodic trends in reactivity (§8.5, pp.286b-293). Ch.9: Half-reactions, activity series, balancing redox reactions, serial dilution. Ch.10: Calorimetric corrections, lattice energy, Born-Haber cycle.

Information provided (see below): Periodic table, solubility rules, 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. Be sure you know how to determine an empirical formula from percent composition data.

Remember to keep practicing and developing the use of the periodic table as an invaluable tool for retrieval of a wealth of chemical information. 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 8 – Chemical Reactions

- Balance simple chemical reactions and understand what the coefficients mean
  §8.1, pp.269-272; Worked Examples 8.1 and 8.2 and Practice Problems A and B for both; Section Review 8.1.1-8.1.4 (p.274); End-of-chapter problems 8.6, 8.7, 8.9 (pp.295-296).
- Determine an empirical formula from combustion analysis data
  §8.2, pp.275-277t; Worked Example 8.4 along with Practice Problems A and B; Section Review 8.2.1-8.1.3 (p.277); End-of-chapter problems 8.16-8.21 (p.296).
- Use stoichiometric relationships in calculations
  §8.3, pp.277-280; Worked Examples 8.5 and 8.6 along with Practice Problems A and B for both; Section Review 8.3.1-8.1.4 (p.280); End-of-chapter problems 8.25-8.37 odd (pp.296-297).
- Determine the limiting reactant, theoretical yield, and percent yield for a reaction
  §8.4, pp.280-286; Worked Examples 8.7 and 8.8 along with Practice Problems A and B for both; Section Review 8.4.1-8.4.5 (p.286); End-of-chapter problems 8.45-8.55 odd (pp.298-299). See also Fig.8.4 (pp.282-283) and Key Skills: Limiting Reactant (pp.304-305).
Chapter 9 – Chemical Reactions in Aqueous Solutions

• Understand what happens when weak, strong, and non-electrolytes dissolve in water

• Predict when a precipitation reaction will occur and what will precipitate

• Write molecular, ionic, and net ionic equations (applies to acid-base reactions as well as precipitation reactions)
  §9.2, pp.312-316; Worked Examples 9.2 and 9.3 along with Practice Problems A and B; Section Review 9.2.1-9.2.5 (p.316); End-of-chapter problems 9.17-9.25 odd (p.350). see also Key Skills: Net Ionic Equations (pp.360-361).

• Identify acids and bases and describe what happens when they are combined

• Determine oxidation numbers and identify redox reactions
  §9.4, pp. 321-325; Worked Example 9.5 along with Practice Problems A and B; Section Review 9.4.1-9.4.3 (p.330); End-of-chapter problems 9.37, 9.44-9.49 (p.351).

• Understand and use the concentration unit of molarity

• Solve stoichiometry problems involving solutions

Chapter 10 – Energy Changes in Chemical Reactions

• Define and give examples of a state function

• Understand the law of conservation of energy
  §10.2, pp.365-368; Worked Example 10.1 along with Practice Problems A and B; Section Review 10.2.1, 10.2.2 (p.368); End-of-chapter problems 10.8, 10.9, 10.11, 10.13 (p.399).

• Understand the difference between change in enthalpy and change in internal energy
  Distinguish these as \(\Delta U = q_V\), measured in constant-volume (“bomb”) calorimetry; \(\Delta H = q_P\), measured in constant-pressure (“coffee-cup”) calorimetry (see §10.3, pp.370-371 and §10.4).

• Relate heat and temperature through heat capacity or specific heat capacity
Chapter 10 learning goals (cont’d.)

- Use calorimetry data to calculate thermal properties
  §10.4, pp.373b-383; Worked Examples 10.4, 10.5 and Practice Problems A and B for both; Section Review 10.4.2-10.4.4 (p.383); End-of-chapter problems 10.26, 10.27, 10.31, 10.35 (pp.400-401). See also Fig.10.9 (pp.376-377) and Fig.10.10 (pp.380-381).

- Understand and apply Hess’s Law

- Define and use standard enthalpies of formation to determine enthalpy changes for a reaction
  §10.5, §10.6, pp.383-387; Worked Examples 10.7, 10.8, 10.9 and Practice Problems A and B for each; Section Review 10.5.1, 10.5.2 (p.385); Section Review 10.6.1-10.6.3 (p.388); End-of-chapter problems 10.40, 10.41, 10.43, 10.45-10.49, 10.57-10.65 odd (pp.401-403). Note that these learning goals rely on the correct interpretation, writing, and manipulation of thermochemical equations (§10.3, p.372).

- Understand what bond enthalpy refers to, understand their relative values, and use them to estimate enthalpy changes for a reaction
  §10.7, pp.388-391; Worked Example 10.10 and Practice problems A and B; Section Review 10.7.1-10.7.4 (p.391); End-of-chapter problems 10.66, 10.67-10.73 odd (p.403).
### Table 9.2 Solubility Guidelines: Soluble Compounds

<table>
<thead>
<tr>
<th>Water-soluble compounds</th>
<th>Insoluble exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing an alkali metal cation (Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺) or the ammonium ion (NH₄⁺)</td>
<td></td>
</tr>
<tr>
<td>Compounds containing the nitrate ion (NO₃⁻), acetate ion (C₂H₃O₂⁻) or chlorate ion (ClO₃⁻)</td>
<td></td>
</tr>
<tr>
<td>Compounds containing the chloride ion (Cl⁻), bromide ion (Br⁻), or iodide ion (I⁻)</td>
<td>Compounds containing Ag⁺, Hg₂⁺, or Pb₂⁺</td>
</tr>
<tr>
<td>Compounds containing the sulfate ion (SO₄²⁻)</td>
<td>Compounds containing Ag⁺, Hg₂⁺, Pb⁺²⁺, Ca⁺²⁺, Sr⁺²⁺, or Ba⁺²⁺</td>
</tr>
</tbody>
</table>

### Table 9.3 Solubility Guidelines: Insoluble Compounds

<table>
<thead>
<tr>
<th>Water-insoluble compounds</th>
<th>Soluble exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing the carbonate ion (CO₃²⁻), phosphate ion (PO₄³⁻), chromate ion (CrO₄²⁻), or sulfide ion (S²⁻)</td>
<td>Compounds containing Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, or NH₄⁺</td>
</tr>
<tr>
<td>Compounds containing the hydroxide ion (OH⁻)</td>
<td>Compounds containing Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, or Ba⁺²⁺</td>
</tr>
</tbody>
</table>

### Selected thermodynamic data (for 25 °C)

<table>
<thead>
<tr>
<th>Substance</th>
<th>∆H°₉₀ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon</td>
<td></td>
</tr>
<tr>
<td>C(s, graphite)</td>
<td>0</td>
</tr>
<tr>
<td>CH₄(g)</td>
<td>−74.6</td>
</tr>
<tr>
<td>C₆H₆(l)</td>
<td>49.1</td>
</tr>
<tr>
<td>CO(g)</td>
<td>−110.5</td>
</tr>
<tr>
<td>CO₂(g)</td>
<td>−393.5</td>
</tr>
<tr>
<td>Hydrogen</td>
<td></td>
</tr>
<tr>
<td>H⁺(aq)</td>
<td>0</td>
</tr>
<tr>
<td>Oxygen</td>
<td></td>
</tr>
<tr>
<td>O₂(g)</td>
<td>0</td>
</tr>
<tr>
<td>H₂O(l)</td>
<td>−285.8</td>
</tr>
</tbody>
</table>
| H₂O(g)          | −241.8
Sample exam questions

The sample questions have appeared on previous exams and are meant to be representative of the kinds of questions to be expected, not necessarily the weight or distribution of topics that will appear on the actual exam.

1. Balance the following chemical equations

(a) \[ ___ \text{Cr}(s) + ___ \text{S}_8(s) \rightarrow ___ \text{Cr}_2\text{S}_3(s) \]

(b) \[ ___ \text{C}_6\text{H}_6(l) + ___ \text{O}_2(g) \rightarrow ___ \text{CO}_2(g) + ___ \text{H}_2\text{O}(g) \]

2. Dimethylhydrazine is a carbon-hydrogen-nitrogen compound used in rocket fuels. When burned completely, a 0.312 g sample yields 0.458 g CO\text{2} and 0.374 g H\text{2}O. The nitrogen content of the sample was converted to 0.145 g of N\text{2}. What is the empirical formula of dimethylhydrazine?

3. Given the following reaction,

\[ 3 \text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2 \text{HNO}_3(aq) + \text{NO}(g) \]

perform the following conversions:

(a) 6.25 mol NO\text{2}(g) to mol HNO\text{3}(aq) \hspace{1cm} (b) 125.0 g NO(g) to g NO\text{2}(g)
4. Consider the following reaction

\[ 6 \text{ClO}_2 + 3 \text{H}_2\text{O} \rightarrow 5 \text{HClO}_3 + \text{HCl} \]

When 50.0 g chlorine dioxide and 25.0 g water are initially present as reactants, the resulting amount of HClO₃ is 48.8 g. Given this information, answer the following:
(a) What is the limiting reactant? (b) Predict the theoretical yield of HClO₃ and calculate the percent yield.

5. Classify each of the following as either a strong electrolyte, weak electrolyte, or nonelectrolyte:
   (a) HClO
   (b) HClO₄
   (c) Mg(OH)₂
   (d) CH₃CH₂OH

6. For each of the following pairs of reactants, use solubility rules to predict whether or not a precipitation reaction occurs. If there is no reaction, indicate this by writing N.R. If a precipitation reaction occurs, write the correct formula and name of the precipitating compound. Note: You do not need to write a chemical equation, just the correct formula and name of any precipitant.
   (a) NaCl(\text{aq}), \text{Ca(NO}_3)_2(\text{aq})
   (b) HNO₃(\text{aq}), \text{K}_2\text{CO}_₃(\text{aq})
   (c) Na₂SO₄(\text{aq}), \text{BaCl}_2(\text{aq})
7. In the following chemical equations, identify the Brønsted-Lowry acids and the Brønsted-Lowry bases, and indicate the conjugate acid/base pairs:

(a) \[ \text{HA}(aq) + \text{B}(aq) \rightleftharpoons \text{A}^- (aq) + \text{BH}^+(aq) \]

(b) \[ \text{H}_2\text{SO}_4(aq) + \text{NH}_3(aq) \rightleftharpoons \text{NH}_4^+(aq) + \text{HSO}_4^-(aq) \]

(c) \[ \text{C}_2\text{H}^- (g) + \text{NH}_3(g) \rightleftharpoons \text{C}_2\text{H}_2(g) + \text{NH}_2^-(g) \]

8. Assign oxidation numbers for the following species:

(i) \( \text{SF}_4 \) 

(ii) \( \text{AsO}_3^{3-} \) 

(iii) \( \text{LiH} \)

<table>
<thead>
<tr>
<th>Species</th>
<th>Ox # S</th>
<th>Ox # F</th>
<th>Ox # As</th>
<th>Ox # O</th>
<th>Ox # Li</th>
<th>Ox # H</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{SF}_4 )</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
</tr>
<tr>
<td>( \text{AsO}_3^{3-} )</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
</tr>
<tr>
<td>( \text{LiH} )</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
<td>____</td>
</tr>
</tbody>
</table>

9. Classify each of the following reactions as precipitation, acid-base, or oxidation-reduction reactions:

(a) \[ \text{Ca(OH)}_2(s) + 2 \text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) \]

(b) \[ \text{SiCl}_4(l) + 2 \text{Mg}(s) \rightarrow 2 \text{MgCl}_2(s) + \text{Si}(s) \]

(c) \[ \text{MgSO}_4(aq) + \text{Na}_2\text{S}(aq) \rightarrow \text{MgS}(s) + \text{Na}_2\text{SO}_4(aq) \]

10. (a) Calculate the molarity of 4.404 g of ammonium sulfate in 250 mL of solution.

(b) Write the molecular (formula) equation, full ionic, and net ionic equation for the reaction in solution between ammonium sulfate and lead (II) nitrate.
Question 10, (cont’d.)
(c) If the lead nitrate solution is at 0.200 M, what volume of the solution is required to react completely with the ammonium sulfate solution specified in part (a)?

11. A stock solution of hydrochloric acid with concentration 12.1 M is to be used to make a dilution to 0.150 M. 
(a) Write the formula for hydrochloric acid. (b) If 500 mL of 0.150 M solution is to be made, what volume of stock solution should be used?

12. When 1 mol of a fuel is burned at constant pressure, it produces 5250 kJ of heat and does 35 kJ of work. Give the sign of \( q \) and \( w \) and determine \( \Delta U \) and \( \Delta H \) for the combustion of the fuel.

13. An experiment is performed in which 50.0 mL of water at 20.00°C is mixed with 50.0 mL of an unknown liquid at 40°C. The final temperature of the mixture is 23.50°C. Using the values 0.9971 g/cm³ and 4.184 J/g°C for the density and specific heat of water, respectively, calculate the specific heat of the unknown liquid if its density is 0.753 g/mL.
14. Use the reactions given below to determine the enthalpy change of the following reaction:

\[
\text{CaO}(s) + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) \quad \Delta H_{\text{rxn}} = ?
\]

\[
\text{Ca}(s) + \text{CO}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{CaCO}_3(s) \quad \Delta H = -812.8 \text{ kJ}
\]

\[
2 \text{ Ca}(s) + \text{O}_2(g) \rightarrow 2 \text{ CaO}(s) \quad \Delta H = -1269.8 \text{ kJ}
\]

15. Using the thermodynamic data provided (see page 4), determine \(\Delta H^\circ_{\text{rxn}}\) for the reaction represented by the balanced equation from Question 1(b).

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

(a) Pressure is a state function.  True ____  False____

(b) Work is a state function.  True ____  False____

(c) \(\Delta U\) for a process equals \(\Delta H\), as long as pressure is held constant.  True ____  False____

(d) Specific heat capacity is an extensive quantity.  True ____  False____
17. For the reaction given below, identify the bonds broken and the bonds formed and from the bond enthalpy data given, estimate the $\Delta H_{\text{rxn}}$.

<table>
<thead>
<tr>
<th>Bond</th>
<th>H</th>
<th>C</th>
<th>O</th>
<th>Cl</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>436</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td>414</td>
<td>347</td>
<td></td>
<td></td>
</tr>
<tr>
<td>O</td>
<td>460</td>
<td>351</td>
<td>142</td>
<td></td>
</tr>
<tr>
<td>Cl</td>
<td>432</td>
<td>339</td>
<td>203</td>
<td>243</td>
</tr>
</tbody>
</table>

$\text{CH}_3\text{OH(g)} + \text{Cl}_2(g) \rightarrow \text{CH}_2\text{Cl}_2(g) + \text{H}_2\text{O(g)}$