Chemistry
STUDY GUIDE #5

Updated 1/8/2011

Chemical Reactions and Quantities

A student who completes this unit should be able to do all of the following:

1) Define a chemical reaction
   A rearrangement of atoms in which compounds may break down and new compounds may form

2) Identify evidence for a chemical reaction
   Heat is either consumed (endothermic process) or released (exothermic process)
   A gas may form (bubbling, new odors, etc.)
   A solid may form (precipitation)
   Colors, odors, and textures may change

3) Write a balanced chemical equation
   1. Make sure all your formulas are correct; you may not change them
   2. Remember the elements that form diatomic molecules: H₂ N₂ O₂ F₂ Cl₂ Br₂ I₂
   3. An element that occurs uncombined should be balanced after everything else
   4. Keep polyatomic ions together whenever possible; if OH occurs, rewrite water as HOH
   5. Balance ions or atoms one at a time, by trial and error
   6. Once several things are in balance, you must keep them in balance; if you multiply
      one of them, you need to multiply them all by the same number.
   7. If you need ½ or 1½ or 2½ or 3½ or 4½ of something,
      then multiply everything by 2. This occurs mostly with O₂.
   8. When you are finished, check everything.

   Things that react are called “reactants” (written on the left)
   Things produced are called “products” (written on the right)

   Optional symbols: (s) solid  (l) liquid  (g) gas  (aq) aqueous (dissolved in water)

4) Explain Oxidation and Reduction:
   Oxidation: Loss of Electrons is Oxidation (remember: LEO)
   Reduction: Gain of Electrons is Reduction (remember that he says GER)

5) Explain four basic types of reactions and identify examples of each:
   a) Synthesis  A + B -- > AB  usually involves oxidation & reduction
   b) Decomposition  AB -- > A + B  usually involves oxidation & reduction
   c) Single replacement  A + BC -- > AC + B  always involves oxidation & reduction
   d) Double replacement  AB + CD -- > AD + CB “changing partners”

   Usually double replacement reactions are reversible, unless:
   i) an insoluble precipitate forms
   ii) a covalent compound (such as water) forms
   iii) a gas forms and escapes
   iv) a product breaks down

   NOTE: Single and double replacement are sometimes called single and double displacement

6) Predict and balance various reactions:
   a) Precipitations: follow a table of solubility rules; AgCl and BaSO₄ are commonly used as examples
   b) Acid-base neutralizations: double replacement reactions in which water is formed
   c) Oxidation and Reduction (remember that LEO says GER)
   d) Combustion reactions: atmospheric oxygen combines with other elements

   CONTINUED on next page
7) Calculate molecular weights and percentage compositions
   a) Multiply each atomic weight by the number of atoms of that type in the formula
   b) Add up the resulting weights for all elements present; total is called Molecular Weight (MW)
   c) Divide the weight of each element by the total (x 100%) to find percentage composition

8) Determine empirical formulas and molecular formulas from composition by mass:
   a) Divide the mass or percentage of each element by its atomic weight
   b) Divide the results by the lowest value in the column
   c) If the results are close to integers, use those integers;
     if not, try multiplying everything by 2 (or by 3, 4, etc.) to get all results close to integer values
   d) The resulting Empirical Formula is the simplest formula consistent with the given composition
   e) The correct Molecular Formula is either the empirical formula or a multiple of the empirical formula;
     you need additional information (such as a molecular weight) to determine the molecular formula

9) Perform calculations involving moles:
   a) One mole = the molecular weight written in grams (also called “molar mass”)
   b) Use the molar mass as a conversion factor: from moles to grams, or from grams to moles

10) Use Avogadro’s law:
    a) One mole of any compound contains \(6.022 \times 10^{23}\) molecules per mole

11) Use STOICHIOMETRY to calculate reaction quantities:
    a) Start with a balanced equation (nothing else works unless this step is correct first)
      Write down the number of moles below each reactant or product
    b) Calculate all molecular weights, then multiply by coefficients in the equation
      Write down the mass of each reactant or product below its formula
    c) Make sure that all quantities are balanced: total mass on left must equal total mass on right
    d) Set up a proportion to solve for the quantity that you want (call it x or y; use algebra)
      i) Each term will have “problem” quantities (from the problem) on top and
         “theoretical” quantities (from the theoretical equation) on the bottom, in the same units
      ii) Solve for x or y by multiplying by the denominator below x or y

12) Perform calculations involving Limiting reactants:
    a) Use stoichiometry to calculate yield separately from each reactant;
    b) Select the limiting reactant as the one that has the smallest yield

13) Perform calculations involving Percentage yield
    \[100\% \times \frac{\text{Actual yield}}{\text{Theoretical yield}} = \text{percent yield}\]
    Note: product must be pure. Percent yield above 100% means product is not pure.

TEXT CHAPTERS: Ch. 8 and Ch. 9
MASSACHUSETTS CURRICULUM FRAMEWORKS, content standard #5

STUDY EVERY NIGHT. COME TO CLASS PREPARED.
ASK FOR HELP IF YOU NEED HELP.
YOU CAN DO WELL IF YOU WANT TO.