CHEMISTRY

STUDY GUIDE FOR MODULES 7C through 8D

HOW TO USE THIS STUDY GUIDE:

This is only a <u>summary</u> of the important points of each chapter.
Use each point to remind you of other concepts in each section. Make sure you <u>understand</u> all terms.
Anything marked with an asterisk (*) involves equations that you should know how to use.
Give yourself practice in these, using old homeworks and quizzes.

Module 7C. Kinetics -

<u>Collision theory</u> refers to the assumption that reactions happen when particles collide with sufficient energy and in proper orientation. Rates are explained with reference to this theory.

* First-order reactions: rate is proportional to the concentration of a reactant: rate = $k\,[A]^1$.
  Decay is exponential, $[A] = e^{-kT}$, $[A]/[A_0] = (1/2)^{T/H}$, where $H = t_{½}$ = half-life
* Second-order reactions: rate = $k\,[A]^2$. Each half-life gets longer.

Zero-order reactions: rate is constant, does not depend on [A].

* Method of initial rates:
  Zero order: Doubling or tripling the initial concentration has no effect on the rate.
  First order: Doubling the initial concentration doubles the rate; tripling it triples the rate.
  Second order: Doubling the initial concentration quadruples the rate; tripling it multiplies x9

Catalysts speed up reactions by reducing activation energy, but they do not get used up in reactions.

Module 7D. Equilibrium chemistry -

An equlibrium is a dynamic state in which <u>no net change</u> occurs because opposite rates are balanced

* "Law of mass action": If $a\,A + b\,B \;\backslash\!=\!=\!=\!=\!=\!\backslash\; c\,C + d\,D$ then

$$k_{eq} = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

* Heterogeneous equilibria:
  Include only gases (g) and aqueous solutions (aq); ignore solids (s) and pure liquids (l)
* LeChâtelier's principle: If a change (or a "stress") is made to a system at equilibrium, the equilibrium will shift in whatever direction undoes the change (or reduces the "stress").

Modules 8A-8B. Acids, Bases, and pH -

Arrhenius' theory: acids produce H^+ ions in solution; bases produce OH^- ions in solution.

Brønsted & Lowry: acids are proton (H^+) donors; bases take protons away (recipients).
* Be sure you can pick out proton donors and recipients in any reaction.

Lewis: bases donate lone pairs of electrons; acids accept lone pairs and bond with them.

Anything that can act as either an acid or a base is called <u>amphiprotic</u>. Water is amphiprotic:

$H_2O + H_2O$ ----> $H_3O^+ + OH^-$ (proton donor becomes OH^-; recipient becomes H_3O^+)

*Review all past quizes and homeworks; make sure you can do problems using these equations.

CONTINUED ON THE REVERSE SIDE

Modules 8A-8B. Acids, Bases, and pH (continued)**-**

A strong acid ionizes completely in solution. HCl, HBr, HNO_3, and H_2SO_4 are strong acids.

Most other acids are weak acids; they ionize only partially in solution.

Similarly, strong bases (like NaOH and KOH) ionize completely, weak bases only partly.

In aqueous solutions: acids and bases neutralize each other to produce water and a salt.

* $pH = - \log[H^+]$; $pH = 7$ if $[H^+] = 10^{-7}$ (neutral)

Acidic solutions have $pH < 7$; basic solutions have $pH > 7$.

$pOH = - \log[OH^-]$; $k_w = [H^+][OH^-] = 10^{-14}$; so $pH + pOH = 14$

Buffers are solutions that resist a change in pH.

Most buffers are a combination of a weak acid and its congugate base.

Modules 8C-8D. Oxidation and Reduction -

* Oxidation numbers are conventional quantities to describe oxidation state (pretended charge):

Uncombined elements have an oxidation number of 0.

Charged ions (incl. polyatomic) have an oxidation number equal to their charge.

Oxygen in compounds has an oxidation number of -2. Halogens are usually -1.

Hydrogen and group I metals usually have oxidation numbers of +1.

Most other oxidation numbers can be figured by making molecules add up to zero.

* Oxidation is an increase in oxidation number, or a loss of electrons.

* Reduction is a reduction (decrease) in oxidation number, or a gain of electrons.

LEO says GER: Loss of Electrons is Oxidation; Gain of Electrons is Reduction.

"REDOX" (reduction & oxidation) reactions always have electrons lost = electrons gained.

* Balancing redox reactions by the half-reaction method:

Separate oxidation and reduction half-reactions; balance them separately:

Balance metals, halogens, and miscellaneous ions first;

Balance O by adding H_2O as needed.

Balance H by adding H^+ ions as needed. Remember: "O, H, 'charge!' "

Balance charge by adding electrons as needed.

Combine half-reactions together, multiplying by constants to make electrons gained = lost

"Disproportionation" means that the same substance is both oxidized and reduced.

An oxidizing agent always oxidizes something else; it undergoes reduction.

A reducing agent always reduces something else; it undergoes oxidation.

Electrochemistry: a Voltaic cell converts chemical energy into electrical energy.

ANode is always the site of OXidation (remember: AN OX)

CAThode is always the site of REDuction (remember: RED CAT)

Electrolysis: passing electricity through a chemical solution results in oxidation & reduction:

OXidation at the ANode and REDuction at the CAThode.

* Be sure you can write oxidation & reduction half-reactions at both anode and cathode.

*Review all past quizes and homeworks; make sure you can do problems using these equations.