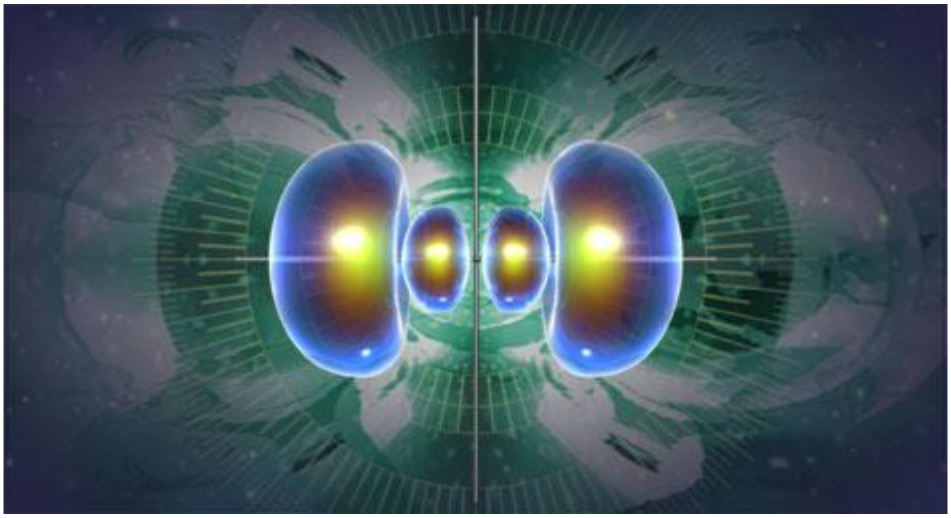

Ms. Cecilia M. Fernández
Honors Chemistry 531

FINAL EXAMINATION REVIEW

General Information:

→*The Final Examination will cover Chapters 4s4&20s2,5-6,8-10 plus Chapter 15*
→*You will be provided with a periodic table.*
→*You will be allowed to use a calculator*
→ *All questions are to be answer on a Scantron.*

Chapter 4s4 & 20s2: Oxidation Reduction

Redox Reactions

LEO – losing electrons **oxidation**
GER – gaining electrons **reduction**

Oxidation States: apparent charge of an atom.

Rules for assigning oxidation states:

1. **Elemental form:** oxidation state = 0.
2. **Monatomic ion:** oxidation state = charge of ion.
3. **Specific elements:**
   a) **Fluorine**: oxidation state = -1.

  b) **Oxygen**: oxidation state = -2, except peroxide = -1.
  c) **Hydrogen**: combined with metals = -1
     combined with nonmetals = +1

4. **If no other rule applies, assign oxidation state equal to the charge of the ion for the last element.**
5. **The sum of the oxidation states of all atoms in a formula equal the charge of the formula.**

Oxidizing agents – substance that contains the element that gets reduced.
Reducing agents – substance that contains the element that gets oxidized.

Balancing net ionc redox reactions

1. **Split into half reactions**
2. **Balance each half reaction:**
   - **a)** balance every element that is not **H** or **O**
   - **b)** balance **O** by adding H_2O.
   - **c)** balance **H** by adding H^+.
   - **d)** balance charge by adding e^-.
3. **Make sure you have the same number of electrons on both half reactions and they are on opposite sides.**
4. **Add the two half reactions.**
5. **If it is in basic solution, add OH^- to both sides, and simplify the water.**

Examples:

1. Write the balanced net ionic equations for the reactions that occur when each of the following reactants are mixed:

   a) Cu (s) + HCl (aq)

2. Balance the following ionic redox reactions, label the oxidizing agent and the reducing agent:

   a) Zn (s) + NO_3^- (aq) $\rightarrow$ Zn^{+2} (aq) +N_2O (g) (acidic solution)

   b) $Fe(OH)_2$ (s) + CrO_4^{2-} (aq) $\rightarrow$ Fe_2O_3 (s) + $Cr(OH)_4^-$ (aq) (basic solution)

33. Determine the oxidation number in each of the following:

a) BrO_3^-
b) $C_2O_4^{2-}$
c) H_4SiO_4
d) SF_6
e) ICl_3

Chapter 5 – Thermochemistry

Kinetic vs potential
Joule = kgm²/s² (kgm^2/s^2)
universe = system + surroundings
open vs. closed vs. isolated
work – force times displacement
enthalpy – heat measured at constant pressure
exothermic vs endothermic
state function – the change does not depend on reaction path.
heat capacity – amount of energy required to increase the temperature of a sample of matter by 1ºC
specific heat – amount of heat required to increase 1g of matter by 1ºC.
enthalpy of formation –enthalpy change for the formation of a substance from its elements in their most common form.

$q = mC\Delta T$

$C_{(H2O)} = 4.184$ **J/gK**

$\Delta H_f = \Sigma\Delta H_f$ **products** $- \Sigma\Delta H_f$ **reactants**

Chapter 10: Gases

Terms:
Volume (L)
Pressure (atm, kPa, mmHg)
Partial Pressure/Atmospheric Pressure
Temperature (K = °C + 273)
Ideal Gas
Diffusion/Effusion
Dalton's Law of Partial Pressures
STP
Mole Fraction

Kinetic-Molecular Theory
Root-mean-square speed

Formulas:

Boyle's Law $P_1V_1 = P_2V_2$

Charles' Law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Gay-Lussac's Law $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Avogadro's Law $\frac{V_1}{n_1} = \frac{V_2}{n_1}$

The Combined Gas Law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

The Ideal Gas Law $\frac{P_1V_1}{T_1n_1} = \frac{P_2V_2}{T_2n_1}$

The Universal Gas Law $PV = nRT$

Dalton's Law of partial pressures $P_t=P_A + P_B + P_C +\ldots..$

Graham's Law of Effusion $\frac{r_A}{R_B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$

Examples:

1. An ideal gas has a volume of 20 L at a certain pressure. By what factor would the volume need to change to halve the pressure?

2. 1.00 mole of oxygen gas occupies 25 L at 300 K. What is its pressure in atmospheres? (R = 0.0821 L·atm/K·mol)

3. A sample of hydrogen gas occupies 50 mL at 20°C and 100 Torr. If the volume of the gas is 300 mL at 30 Torr, what is the Celsius temperature?

4. List some properties of gases.

5. If the partial pressures of oxygen and carbon dioxide are 245 Torr and 35 Torr, respectively, and the atmospheric pressure is 760 Torr, what is the partial pressure of nitrogen gas if it makes up the remainder of the atmosphere?

Chapter 6: Models of the Atom

Terms:
Orbitals
Energy Levels
Quantum Numbers
Electromagnetic radiation
Uncertainty Principle
Plank's Constant
Photoelectric Effect
Line spectrum
Ground state vs excited state
Matter waves
Shell
Subshell
Degenerate
Pauli Exclusion Principle
Electron Configuration
Hund's Rule
Valence electrons vs core electrons

Formulas:
Speed of light $c=\lambda\upsilon$

Energy of a photon $E_p = h\upsilon$

$E_p = hc/\lambda$

Photoelectric Effect $E_p = E_b + KE_e$

Matter waves $\lambda = h/mv$

EXAMPLES:

1. Give the maximum number of electrons for the following orbitals:
   s p d f

2. Give the maximum number of electrons for the following energy levels:
   1st 2nd 3rd 4th

3. Write the electron configurations for:

   Br

   N^{3-}

   Pt

   Cu^{+}

Chapter 8, 9: Chemical Bonding, VSEPR

Terms:

Valence electrons
Lewis (electron dot) structures
Octet rule
Ionic bond
Metallic bond
Covalent bond:
 Single bond, Double bond, Triple bond
 Coordinate Covalent bond
Lone pair or Unshared pair
Polar vs. Non-polar
Dipole moment
Formal charge
Resonance structures
Bond angles
Electron domains
VSEPR theory: electron pair geometry and molecular shape
 Linear, Bent, Trigonal Planar, Tetrahedral, Trigonal Bypyramidal, Octahedral

Overlap
Hybridization
Sigma vs pi bonds
Delocalized electrons
Bond order

Examples:

1. Draw Lewis structures for these atoms and ions:

P C Br^- Ag^+ O^{2-}

2. Which of the following exhibit ionic bonding?

NaCl H_2O SiO_2 Fe KNO_3 N_2

3. Draw Lewis structures for the following molecules and ions:

H_2O N_2 CO_2

SO_4^{2-} $AlCl_3$

4. Name the geometry and shape for each of the molecules you just drew.

5. Circle any polar molecules:

H_2O CO O_2 NH_4^+ PCl_5 Au

Chapter 15: Equilibrium

Terms:
Equilibrium
Collision Theory
Exothermic vs. Endothermic reaction profiles
Heat of Reaction
Activation energy
Catalyst
Homogeneous vs heterogeneous equilibrium
Reaction quotient (Q)
Le Chatelier's Principle

Formulas:

$$K_c = \frac{[\text{products}]}{[\text{reactants}]}$$

$$K_p = \frac{P_{(\text{products})}}{P_{(\text{reactants})}}$$

$$K_p = K_c(RT)^{\Delta n}$$

Examples:

1. Which direction will the equilibrium shift if the following stresses are applied to the reaction:

   $2C_{(s)} \quad + \quad H_{2(g)} \quad + \quad O_{2(g)} \quad \rightarrow \quad 2CO_{(g)} \quad + \quad 2H^{+}_{(aq)}$

   a) Increased pressure in the reaction vessel.

   b) CO removed from the reaction vessel.

   c) Water added to the reaction vessel.

   d) If the reaction is endothermic, what will happen if the reaction vessel is heated?

   e) pH decreased.

   f) NaOH added.

   g) Solid C added.

2. Write the equilibrium constant expression for the reaction, then solve for Keq:

$Mg_{(s)}$ +	$2HCl_{(aq)}$	$\rightarrow$	$H_{2(g)}$ +	$MgCl_{2(aq)}$
	4.0 M		**6.0 M**	**5.0 M**