AP CHEMISTRY
REVIEW

AP EXAM INFORMATION

Exam structure: 3 HOURS

Section I – 50% of your score - 90 minutes

60 multiple choice questions (four answer options), no calculator allowed

Section II – 50% of your score

Graphing calculator allowed – 90 minutes

three long free response (about 20 minutes each)

four short free response (about 7 minutes each)

** During each section you will be supplied with a periodic table and a formula and constants chart.

AP CHEMISTRY REVIEW: BIG IDEA # 1
PROPERTIES OF MATTER
Molecules and Atoms

Must Know:

Differences between elements, compounds and mixtures
Explain that separation of mixtures is based on their physical properties
Use SI units for measurement
Apply dimensional analysis and significant figures to calculations
Use the mole as a quantitative model for chemical composition
Interconvert moles, mass, number of particles and volume of a gas
Use calculations of mass data to determine the identity or purity of various substances
Calculate the percentage composition of a compound
Calculate the empirical and molecular formulas of a compound and of a hydrate from combustion and decomposition data
Apply mathematical calculations to mass data to infer the identity of a substance and/or its purity
Calculate wave properties including frequency, wavelength and energy of a photon
Write ground state electron configurations of atoms using the periodic table as a guide
Explain the photoelectric effect and photoelectron spectroscopy (PES)
Describe electron structure using photoelectron spectroscopy, ionization energy data, and Coulomb's Law
Analyze data that relate ionization energies to electron configurations
Explain electron configurations using Coulomb's Law
Identify and distinguish between paramagnetic and diamagnetic electron configurations
Explain how properties of elements vary across the periodic table using Coulomb's law and effective nuclear charge
Explain role that electron configurations play in determining periodic properties
Trends in atomic radius, ionic radius, and ionization energy and rationalize them by applying the ideas of Coulomb's law, effective nuclear charge and the shielding effect.
Identify and explain the anomalies in the trends of first ionization energy
Explain how the technique of chromatography uses intermolecular attractions to separate mixtures

_____ 1. A 4.5g sample of which of the following would have the greatest mass percent of oxygen?

a. Na_2O (molar mass = 62 g/mol)
b. b. Li_2O (molar mass = 30 g/mol)
c. c. MgO molar mass = 40 g/mol
d. d. SrO (molar mass = 104 g/mol

Empirical Formula: Molecular Formula:

% to mass, mass to moles, divide by the smallest and multiple to get whole.

_____2. A compound is determined to be 14 g of nitrogen and 32 g of oxygen. The empirical formula of the compound is

a. NO bN$_2$O c. NO_2 d. NO_3

When heating a hydrate, or ANYTHING multiple times…. heat several times to ENSURE WATER is DRIVEN OFF.

_____ 3. The mass percent of oxygen in pure cluose is 53.3 g percent. A chemist analyzes a sample of gluces that contains impurities and determines that the mass percent of oxygen in is 49.7 percent. Which of the following impurities could account for the low mass percent of oxygen in this sample?

a. $C_{20}H_{42}$ b. $C_{10}H_5O_5$ c. $C_6H_{12}O_6$ d. $C_{12}H_{22}O_{11}$

1 mol = 6.02 E 23 particles   1 mole = molar mass   1 mol = 22.4 L of gas at STP

_____ 4. How many moles of carbon are in 88 grams of propane C_3H_8

a. 2   b. 16   c. 6   d. 98

_____ 5. A compound contains 1.10 mol of K, 0.55 mol of Te, and 1.65 mol of O. What is the simplest formula of this compound?

(a) KTeO   (b) KTe_2O (c) K_2TeO_3 (d) K_2TeO_6 (e) K_4TeO_6

_____ 6. The simplest formula for an oxide of element X (MM = 76.0) that is 24.0 percent oxygen by weight is

(a) X_2O (b) XO   (c) XO_2 (d) X_2O_3 (e) X_2O_5

_____ 7. A 27.0-gram sample of an unknown hydrocarbon was burned in excess oxygen to form 88.0 grams of carbon dioxide and 27.0 grams of water. What is a possible molecular formula of the hydrocarbon?

(a) CH_4 (b) C_2H_2 (c) C_4H_3 (d) C_4H_6 (e) C_4H_{10}

_____ 8. When a hydrate of X_2CO_3 (MM = 153) is heated until all the water is removed, it loses 54 percent of its mass. The formula of the hydrate is

(a) $X_2CO_3 \bullet 10\ H_2O$ (b) $X_2CO_3 \bullet 7\ H_2O$ (c) $X_2CO_3 \bullet 5\ H_2O$ (d) $X_2CO_3 \bullet 3\ H_2O$ (e) $X_2CO_3 \bullet H_2O$

FREE RESPONSE REVIEW: Empirical and molecular formulas

9.) Answer the following questions related to hydrocarbons.

(a) Determine the empirical formula of a hydrocarbon that contains 85.7 percent carbon by mass.

(b) The density of the hydrocarbon in part (a) is 2.0 g L^{-1} at $50^{O}C$ and 0.948 atm.

a. Calculate the molar mass of the hydrocarbon.

b. Determine the molecular formula of the hydrocarbon.

Electronic structure of Atom: Electrons are most stable in their LOWEST ENERGY STATES.

Electron Configurations

	At #	Increasing Energy ⟶ 1s	2s	2p	3s	Electron Configuration
H	1	↑				$1s^1$
He	2	↑↓				$1s^2$
Li	3	↑↓	↑			$1s^22s^1$
Be	4	↑↓	↑↓			$1s^22s^2$
B	5	↑↓	↑↓	↑ _ _		$1s^22s^22p^1$
C	6	↑↓	↑↓	↑ ↑ _		$1s^22s^22p^2$
N	7	↑↓	↑↓	↑ ↑ ↑		$1s^22s^22p^3$
O	8	↑↓	↑↓	↑↓ ↑ ↑		$1s^22s^22p^4$
F	9	↑↓	↑↓	↑↓ ↑↓ ↑		$1s^22s^22p^5$
Ne	10	↑↓	↑↓	↑↓ ↑↓ ↑↓		$1s^22s^22p^6$
Na	11	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑	$1s^22s^22p^63s^1$ ([Ne]$3s^1$)
Mg	12	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	$1s^22s^22p^63s^2$ ([Ne]$3s^2$)

Coulomb's Law – Gives the electric force between two point charges.

$$F = k\frac{q_1 q_2}{r^2}$$ *Inverse Square Law*

k = Coulomb's Constant = 9.0×10^9 Nm^2/C^2

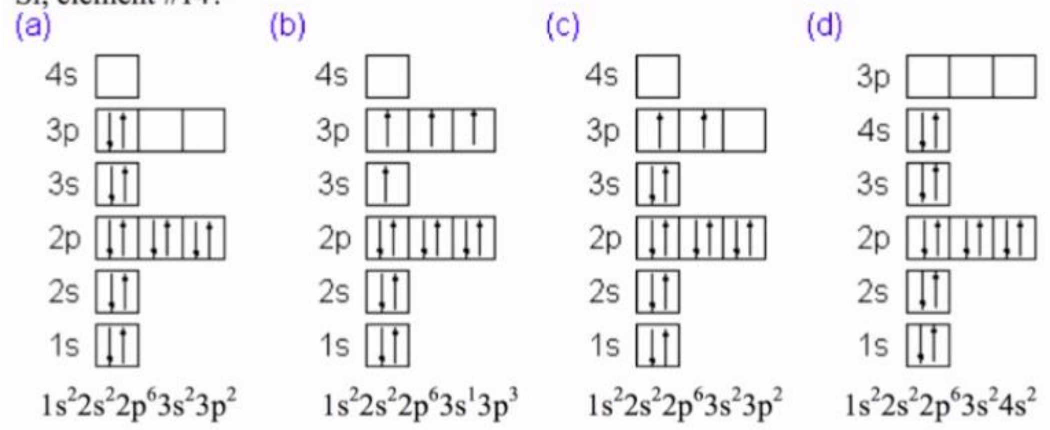
q_1 = charge on mass 1

q_2 = charge on mass 2

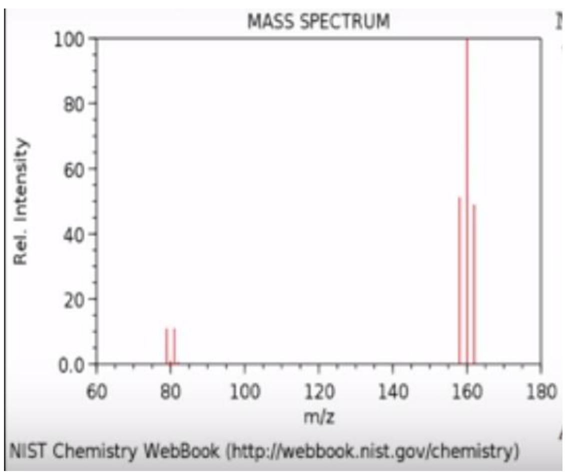
r = the distance between the two charges

The electric force is much stronger than the gravitational force.

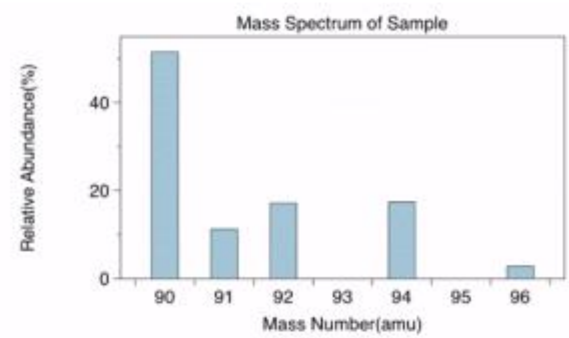
_____ 10. Which of the following electron configurations and orbital diagrams represents Si, element #14?



Mass Spectroscopy



11.



Both Zr and Nb have similar average atomic masses. Above is the mass spectrum of a sample that was believed to have both Zr and Nb. Which of the following statements would be the best explanation for deciding that the sample contained only Zr?

Spectrosopy to measure properties

IR Radiation – Look at bonds within a molecules to see how much they vibrate

UV or X Ray Radiation - this is what we can see
PES
Causes electron transitions

_____ 12. Which one of the following contains a transition paired correctly with an area of high absorption in the electromagnetic spectrum

a. Electronic transistion, microwaves
b. Electronic transitions, infrred radiation
c. Molecular rotation, infrared radiation
d. Molecular vibration, infrared radiation

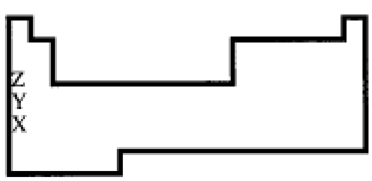
Beer Law – MUST HAVE COLORED SOLUTION

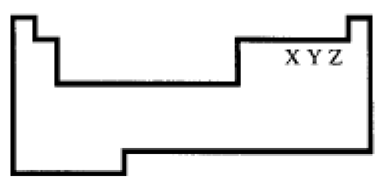
A = abc   abs vs. conc.

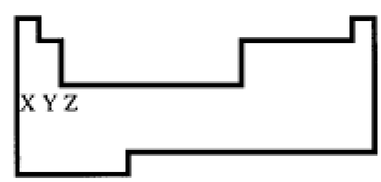
Trends of Periodic Table

Atomic Radii: half the distance between nuclei of identical atoms that are bonded together.
Decreases from Left to Right because the number of electron shells increases but the Zeff increases as the number of Protons increases which pulls the electrons in tighter and therefore the atom is smaller.

Increase from Top to Bottom because of an increased number of electron shells.







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1st Ionization Energy: amount of energy needed to remove the most loosely bound electron from a neutral atom in the gaseous state.
when an atom/ion is smaller (less shells, more Zeff) the nucleus has a greater hold on it's electrons and therefore greater energy is required to remove that electron, therefore higher ionization energy.

2nd Ionization Enrgy: amount of energy required to remove the second most loosely bound electron

3rd Ionization Energy: amount of energy required to remove the third most loosely bound electron.

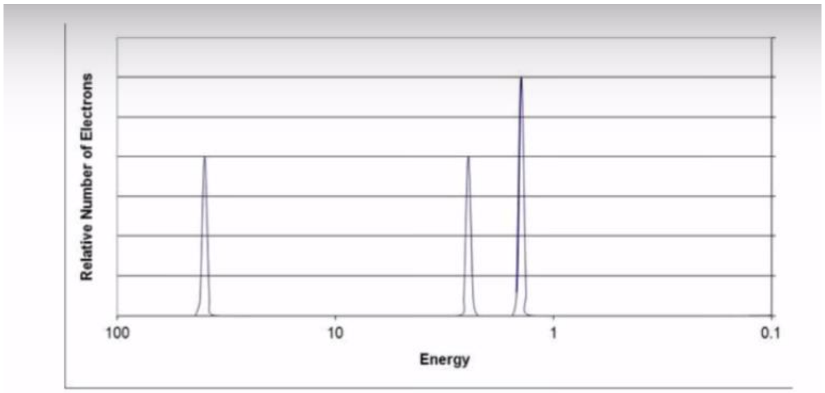
Electronegativity: a measure of its attraction for electrons. "want for more electrons"
when an atom/ion is smaller (less shells, more Zeff) the nucleus has a greater ability to attract electrons and therefore greater ability to pull an additional electron into it's valence shell and therefore higher ionization energy.

Ionic Radius: Gain electrons, Radius increases: **Negative Ions** have a **larger ionic radius** than atomic radius due to greater electron electron repulsions in the valence shell.
Lose electrons, Radius decreases: **Positive Ions** have a **smaller ionic radius** compared to atomic radius due to the loss of the valence shell so the ion has fewer shells and therefore is smaller.

** Shells WINS, then compare Zeff to determine radius of an atom or ion

Photoelectronic Spectra Data (PES)– explains binding energy to the nucleus.
Higher energy means CLOSER TO NUCLEUS, Lower energy means further from nucleus

_____ 13.



Which element could be represented by the complete PES spectrum above?

a. Li
b. B
c. N
d. O

______14. What is the most likely electron configuration for a sodium ion in its ground state?
(A) $1s^2 2s^2 2p^5$ (B) $1s^2 2s^2 2p^6$ (C) $1s^2 2s^2 2p^6 3s^1$ (D) $1s^2 2s^2 2p^5 3s^2$ (E) $1s^2 2s^2 2p^6 3s^2$

______15. Which of the following statements is true regarding sodium and chlorine?
(A) Sodium has a greater electronegativity and a larger first ionization energy.
(B) Sodium has a larger first ionization energy and a larger atomic radius.
(C) Chlorine has a larger atomic radius and a greater electronegativity.
(D) Chlorine has a greater electronegativity and a larger first ionization energy.
(E) Chlorine has a larger atomic radius and a larger first ionization energy.

______16. Which of the following elements is diamagnetic?
(A) H (B) Li (C) Be (D) B (E) C

______17. Which of the following is true of the alkali metal elements?
(A) They usually take +2 oxidation state.
(B) They have oxides that act as acid anhydrides.
(C) They form covalent bonds with oxygen.
(D) They are generally found in nature in compounds.
(E) They have relatively larger first ionization energies.

______18. Which of the following nuclei has 3 more neutrons than protons
(Remember: the number before the symbol indicates atomic mass.)
(A) ^{11}B (B) ^{37}Cl (C) ^{24}Mg (D) ^{70}Ga (E) ^{19}F

______19. Which of the following ions has the smallest ionic radius?
(A) O^{2-} (B) F^- (C) Na^+ (D) Mg^{2+} (E) Al^{3+}

______20. The ionization energies for an element are listed in the table below.

First	Second	Third	Fourth	Fifth
8eV	15eV	80eV	109eV	141eV

Based on the ionization energy table, the element is most likely to be
(A) Sodium (B) Magnesium (C) Aluminum (D) Silicon (E) Phosphorous

______21. A researcher listed the first five ionization energies for a silicon atom in order from first to fifth. Which of the following lists corresponds to the ionization energies for silicon?

(A) 780 kJ, 13,675 kJ, 14,110 kJ, 15,650 kJ, 16,100 kJ
(B) 780 kJ, 1,575 kJ, 14,110 kJ, 15,650 kJ, 16,100 kJ
(C) 780 kJ, 1,575 kJ, 3,220 kJ, 15,650 kJ, 16,100 kJ
(D) 780 kJ, 1,575 kJ, 3,220 kJ, 4,350 kJ, 16,100 kJ
(E) 780 kJ, 1,575 kJ, 3,220 kJ, 4,350 kJ, 5,340 kJ

FREE RESPONSE: Atom and Periodic Table

Problem 22:

Suppose that a stable element with atomic number 119, symbol Q, has been discovered.

(a) Write the ground-state electron configuration for Q, showing only the valence-shell electrons.

(b) Would Q be a metal or a nonmetal? Explain in terms of electron configuration.

(c) On the basis of periodic trends, would Q have the largest atomic radius in its group or would it have the smallest? Explain in terms of electronic structure.

(d) What would be the most likely charge of the Q ion in stable ionic compounds?

(e) Assume that Q reacts to form a carbonate compound.

(i) Write the formula for the compound formed between Q and the carbonate ion, CO_3^{2-}.

(ii) Predict whether or not the compound would be soluble in water. Explain your reasoning.

Problem 23: Explain each of the following in terms of atomic and molecular structures.

a. The first ionization energy for magnesium is greater than the first ionization energy for calcium.

b. The first and second ionization energies for calcium are comparable, but the third ionization energy is much greater.

c. Solid sodium conducts electricity, but solid sodium chloride does not.

d. The first ionization energy for aluminum is lower than the first ionization energy for magnesium.

e. Explain the trend in atomic radius from Al to Mg to Na.

AP CHEMISTRY REVIEW: BIG IDEA # 2
PROPERTIES OF MATTER
Bonding and Phases

Must Know:

Explain how and why potential energy varies with distance between atoms in covalent bonds and intermolecular forces
Explain how temperature relates to molecular motion using particle views
Write both octet and non-octet Lewis structures for atoms, ions and molecules
Predict properties of binary compounds based on their chemical formulas
Provide explanations of bonding properties based on particle views
Predict bonding type in binary compounds based on element positions on the periodic table
Rank bond polarity using electronegativity values of the bonded atoms and/or their positions on the periodic table
Draw resonance structures and interpret the nature of delocalized electrons
Correlate bond multiplicity to bond length and strength
Use formal charge to evaluate the suitability of different octet resonance structures
Distinguish between electron domain geometry and molecular geometry
Apply VSEPR theory to predict geometries of both octet and non-octet Lewis structures of ions and molecules
Predict molecular geometries and bond angles of molecules
Predict sp, sp^2, sp^3 hybridization from the geometries of molecules
Determine the polarity of molecules from their geometries and atomic electronegativities
Visualize covalent bond formation as the overlap of atomic orbitals
Distinguish between sigma and pi bonds
Explain why sigma bonds have larger bond energies than pi bonds
Use Lewis structures to identify formulas that have delocalized electrons
Use bond order to predict relative bond energies
Use kinetic-molecular theory and Maxwell-Boltzmann distribution to explain and make predictions about the macroscopic behavior of the properties of gases: pressure, volume, the number of particles and temperature
Construct particle representations of the gas phase that explain the macroscopic properties of gases
Calculate temperature, pressure and volume from given data for an ideal gas
Interpret the graphical representations of pressure, volume and temperature and explain how absolute zero can be determined experimentally
Explain deviations of real gases from ideal behavior using the structure of atoms and molecules and the forces acting between them
Distinguish among intermolecular forces and explain with examples how they affect the properties of molecules
Compare the macroscopic differences between solids, liquids, and gases using their molecular structures and behaviors and the forces that hold them together
Use particle representations to rationalize the differences between solids, liquids and gases
Use intermolecular forces to explain the properties of liquids such as melting point, boiling point, vapor pressure, viscosity and surface tension.
Use the structures of molecules to predict the types of intermolecular forces that exist between them
Describe the relationships between the structures of polar molecules and their dipole-dipole intermolecular forces.
Use London dispersions to justify properties of nonpolar atoms and molecules
Construct atomic-level visual representations for the structural features of the major classes of solids: ionic, metallic, covalent-network, and molecular
Predict characteristic properties and structures of solids based on their chemical formulas
Distinguish and explain the forces that bind the atoms and molecules of each type of solid
Use 3-D representations and the interaction of particles to explain the common macroscopic properties of each class of solid
Rationalize how the electron-sea model of a metal with its delocalized electrons explains common metallic properties
Describe an alloy and explain how its structure and properties compare to that of a pure metallic solid
Distinguish the structures of substitutional alloys and interstitial alloys
Use chemical formulas to associate and recognize the classification, structure and bonding of solid substances
Design a plan to collect or interpret data to classify a solid substance based on its observable properties.

Make a bond: **release Energy (exothermic)**, bonded atoms have **lower potential energy.** Greater energy released, more stable bond.

$$A + B \rightarrow AB + \text{energy}$$

Break a bond: absorb Energy (endothermic), free atoms have **higher potential energy.**

$$AB + \text{energy} \rightarrow A + B$$

The Octet Rule: Atoms want a complete octet of eight valence electrons in order to be stable. To reach this goal of 8 valence electrons, atoms react by gaining, losing, or sharing electrons. Their goal is to have the **electron configuration of a noble gas.**

Types of Bonds:

Metallic	Ionic	Covalent
Metals **"sea of mobile valence electrons"** Properties: - always good conductors, because of **delocalized valence electrons** - malleable (made into sheets) - ductile (made into wires) - luster (shiny) - transition metals have color in solution	Metal + Nonmetal Strong electrostatic force of attraction between oppositely charged ions **Transfer of electrons** from metal to nonmetal **Electronegativity difference** is **> 1.7** Properties: - brittle and granular due to strong electrostatic forces of attraction between oppositely charged ions - Good conductor as Liquid or (aq) ONLY due to **free moving ions** **-** soluble/ electrolyte - High melting/freezing point - High boiling point - Fast reaction rate	Nonmetal + Nonmetal Attraction of positive nucleus of one atom for the negatively charged electrons of the other atom **Share electrons** To achieve a stable arrangement of electrons Properties**:** -Soft due to weak IMFs - Poor conductors of heat and electricity because no charged mobile particles - Insoluble in water /nonelectrolyte - Low melting and - Low boiling point - Slow reaction rate

Design an experiment to determine if a substance is Ionic or Covalent.

Ionic and Covalent is NOT a binary: in general, as the oxidation state of a metal increases, so does the degree of covalent bonding

VSEPR theory: valence shell electron pair repulsion theory; determines the shapes of molecules

**** Must have available d orbitals in order to expand octet**

OBEY OCTET: 4 electron domains Tetrahedral electronic geometry Hybridization sp^3	
	molecular geometry
central atom NO lone pairs	Tetrahedral
central atom ONE lone pair	Pyramidal
central atom TWO lone pair	bent
central atom THREE lone pair	linear
Only two atoms	linear

VIOLATE OCTET: 5 electron domains Trigonal bipyramidal electronic geometry Hybridization sp^3d	
	molecular geometry
central atom NO lone pairs	Trigonal bipyramidal
central atom ONE lone pair (on equatorial position)	seesaw
central atom TWO lone pair (on equatorial position)	T-shaped
central atom THREE lone pair (on equatorial position)	linear

VIOLATE OCTET: 6 electron domains Octahedral electronic geometry Hybridization sp^3d^2	
	molecular geometry
central atom NO lone pairs	Octahedral
central atom ONE lone pair (on axial position)	Square pyramidal
central atom TWO lone pair (on axial position)	Square planar

Types of Covalent Bonds/molecular bonds:

Nonpolar Bonds: there is an **equal sharing** of electrons, **no pull** on electrons. Formed when atoms in the bond have the same electronegativity. All diatomic molecules have nonpolar **bonds** ($At_2H_2O_2N_2Cl_2Br_2I_2F_2$)

Electronegativity difference between atoms in the bond is between **0 and 0.4**!

Polar Bonds: there is **unequal sharing** of electrons, **pull** on electrons. Found when atoms in the bond have a **different electronegativity**. The element with a **higher electronegativity** has a greater attraction for electrons and ends up with a **partial negative charge**. The other end of the polar covalent bond, with a lower electronegativity acquires a partially positive charge. **Electronegativity difference** is between **0.4 - 1.7**

Types of Molecules:

Nonpolar Molecules:

1- All molecules that are made up of **nonpolar bonds only** are nonpolar molecules.(example all diatomic molecules)
2- Compounds that are symmetrical: have identical parts on each side of it's axis are nonpolar
Memorize: CH_4, CO_2,
3- Have dispersion forces between them and therefore have lower mp, bp, and higher vp

Polar Molecules:

1- molecules made up of **polar bonds** that are asymmetrical: lack identical parts on each side of it's axis
Memorize: H_2O, CH_3Cl , NH_3, H_2S
2- causes molecules to have **dipole-dipole attractions** between them and therefore have higher mp, bp, and lower vp

Single: One bond (pair of electrons) between atoms, **share 2 electrons**, sigma bond
Double: Two bonds (2 pairs of electrons) between atoms, **share 4 electrons**, sigma bond + pi bond
Triple: **Three bonds** (3 pairs of electrons) between atoms, **share 6 electrons**. sigma bond + pi bond + pi bond (shortest due to greater build up of electron density between to two nuclei pulling them in closer and strongest)

Valence Bond Theory: electrons in a covalent bond reside in a region that is an overlap of individual atomic orbitals.
Hybridization: sp, sp^2, sp^3, sp^3d, sp^3d^2

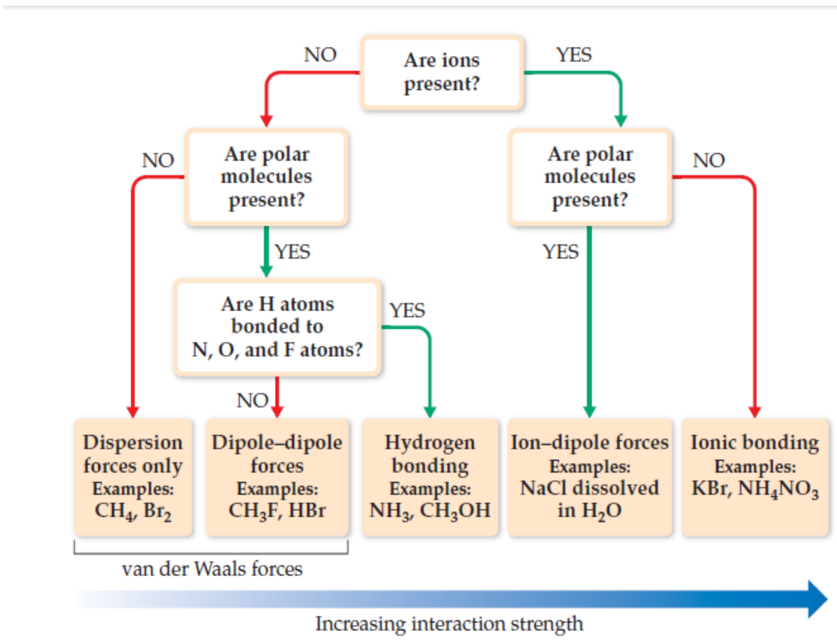
Intermolecular Forces: Attractive forces between molecules that are responsible for the physical properties of molecules: melting point, boiling point, vapor pressure, surface tension.
Stronger IMF means Higher BP, MP, FP and lower VP

Vander walls/London Dispersion Forces: weakest IMF based on # electrons and polarizability, exists between all atoms and molecules

Dipole – Dipole: stronger than LDF, occurs only in polar molecules

Hydrogen Bonds: strongest attractive force, occur only in molecules that have
Hydrogen bonded to F, O, or N.

Keeps DNA together… show how hydrogen bonding exists in a biological system.



Multiple types of intermolecular forces can be at work in a given substance or mixture. In particular, dispersion forces occur in all substances.

GASES: Kinetic Molecular Theory (KMT)

1- Neglect the volume of the molecules

2- Particles are in constant straight line motion and collisions with walls of container causes the pressure exerted by the gas
3- Ignore IMF's of particles (no attractive or repelling forces)
4- KE $_{avg}$ is <u>proportional</u> to the kelvin temperature (heat 'em up, speed 'em up)

Temperature must always be in KELVIN!!

$PV = nRT$

$P_1V_1T_2 = P_2V_2T_1$

Density gas @ STP:

STP: 273 K and 1atm (101.2 kPa , 760. mmHg)

Not @ STP:

Standard Conditions: 25ºC or 298 K, 1 atm

Daltons Law:

"collect over water"

Grahams Law:

Gas Law Graphs:

MULTIPLE CHOICE REVIEW: Gases and Gas Laws

_____24. A gaseous mixture containing 7.0 moles of nitrogen, 2.5 moles of oxygen, and 0.50 mole of helium exerts a total pressure of 0.90 atmospheres. What is the partial pressure of the nitrogen?

(a) 0.13 atm (b) 0.27 atm (c) 0.63 atm (d) 0.90 atm (e) 6.3 atm

_____25. Hydrogen gas is collected over water at 24 °C. The total pressure of the sample is 755 millimeters of mercury. At 24 °C, the vapor pressure of water is 22 millimeters of mercury. What is the partial pressure of the hydrogen gas?

(a) 22 mm Hg (b) 733 mm Hg (c) 755 mm Hg (d) 760 mm Hg (e) 777 mm Hg

_____26. A 2.00-liter sample of nitrogen gas at 27 °C and 600. millimeters of mercury is heated until it occupies a volume of 5.00 liters. If the pressure remains unchanged, the final temperature of the gas is

(a) 68 °C (b) 120 °C (c) 477 °C (d) 677 °C (e) 950. °C

_____27.The density of an unknown gas is 2.00 grams per liter at 3.00 atmospheres pressure and 127 °C. What is the molecular weight of this gas? (R = 0.0821 liter-atm / mole-K)

(a) 254/3 R (b) 188 R (c) 800/3 R (d) 600 R (e) 800 R

_____28. At 20. °C, the vapor pressure of toluene is 25 millimeters of mercury and that of benzene is 75 millimeters of mercury. An ideal solution, equimolar in toluene and benzene, is prepared. At 20. °C, what is the mole fraction of benzene in the vapor in equilibrium with this solution?

(a) 0.25 (b) 0.33 (c) 0.50 (d) 0.75 (e) 0.83

_____29. A sample of 0.010 mole of oxygen gas is confined at 127 °C and 0.80 atmospheres. What would be the pressure of this sample at 27 °C and the same volume?

(a) 0.10 atm (b) 0.20 atm (c) 0.60 atm (d) 0.80 atm (e) 1.1 atm

_____30. A sample of 9.00 grams of aluminum metal is added to an excess of hydrochloric acid. The volume of hydrogen gas produced at standard temperature and pressure is

(a) 22.4 liters (b) 11.2 liters (c) 7.46 liters (d) 5.60 liters (e) 3.74 liters

_____31. A flask contains 0.25 moles of $SO_2(g)$, 0.50 moles of $CH_4(g)$, and 0.50 mole of $O_2(g)$. The total pressure of the gases in the flask is 800 mm Hg. What is the partial pressure of the $SO_2(g)$ in the flask?

(a) 800 mm Hg (b) 600 mm Hg (c) 250 mm Hg (d) 200 mm Hg (e) 160 mm Hg

_____32. $CS_2(l) + 3\ O_2(g) \rightarrow CO_2(g) + 2\ SO_2(g)$

What volume of $O_2(g)$ is required to react with excess CS2(l) to produce 4.0 liters of $CO_2(g)$? (Assume all gases are measured at 0ºC and 1 atm.)

(a) 12 L (b) 22.4 L (c) 1/3 x 22.4 L (d) 2 x 22.4 L (e) 3 x 22.4 L

_____33. A 2 L container will hold about 4 g of which of the following gases at 0ºC and 1 atm?

(a) SO_2 (b) N_2 (c) CO_2 (d) C_4H_8 (e) NH_3

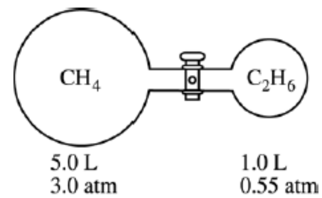
_____ 34. An excess of Mg(s) is added to 100. mL of 0.400 M HCl. At 0ºC and 1 atm pressure, what volume of H_2 gas can be obtained?

(a) 22.4 mL (b) 44.8 mL (c) 224 mL (d) 448 mL (e) 896 mL

FREE RESPONSE REVIEW: Gases and Gas Laws

Problem 35:

(a) Two flasks are connected by a stopcock as shown below. The 5.0 L flask contains CH_4 at a pressure of 3.0 atm, and the 1.0 L flask contains C_2H_6 at a pressure of 0.55 atm. Calculate the total pressure of the system after the stopcock is opened. Assume that the temperature remains constant.



(b) Octane, $C_8H_{18(l)}$, has a density of 0.703 g mL^{-1} at 20ºC. A 255 mL sample of C_8H_{18} reacts completely with excess oxygen as represented by the equation: $2\ C_8H_{18(l)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2\,(g)} + 18\ H_2O_{\,(g)}$

Calculate the total number of moles of gaseous products formed.

Problem 36: $2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$

The mass of an aqueous solution of H_2O_2 is 6.951 g. The H_2O_2 in the solution decomposes completely according to the reaction represented above. The $O_2(g)$ produced is collected in an inverted graduated tube over water at 23.4°C and has a volume of 182.4 mL when the water levels inside and outside of the tube are the same. The atmospheric pressure in the lab is 762.6 torr, and the equilibrium vapor pressure of water at 23.4°C is 21.6 torr.

(a) Calculate the partial pressure, in torr, of $O_2(g)$ in the gas-collection tube.

(b) Calculate the number of moles of $O_2(g)$ produced in the reaction.

(c) Calculate the mass, in grams, of H_2O_2 that decomposed.

(d) Calculate the percent of H_2O_2 , by mass, in the original 6.951 g aqueous sample.

Problem 37:

A rigid 8.20 L flask contains a mixture of 2.50 moles of H_2, 0.500 mole of O_2, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127 ^{o}C.

(a) Calculate the total pressure in the flask.

(b) Calculate the mole fraction of H_2 in the flask.

(c) Calculate the density (in g L^{-1}) of the mixture in the flask.

The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

(d) Give the mole fraction of all species present in the flask at the end of the reaction.

MULTIPLE CHOICE PRACTICE:

Questions 38-41

(A) Metallic Bonding
(B) Network Covalent Bonding
(C) Hydrogen Bonding
(D) Ionic Bonding
(E) London Dispersion Forces

_____ 38. Solids exhibiting this kind of bonding are excellent conductors of heat.
_____ 39. This kind of bonding is the reason that water is more dense than ice.
_____ 40. This kind of bonding exists between atoms with very different electronegativites.
_____ 41. The stability exhibited by diamonds is due to this kind of bonding.

Questions 42-44

(A) CH_4
(B) NH_3
(C) NaCl
(D) N_2
(E) H_2

_____ 42.This substance undergoes ionic bonding.
_____ 43. This molecule contains two pi (π) bonds.
_____ 44. This substance under goes hydrogen bonding.

Questions 45-47

(A) BF_3
(B) CO_2
(C) H_2O
(D) CF_4
(E) PH_4

_____ 45. The central atom in this molecule forms sp^2 hybrid orbitals.
_____ 46. This molecule has a tetrahedral structure.
_____ 47. This molecule has a linear structure.

_____ 48. A liquid whose molecules are held together by which of the following forces would be expected to have the lowest boiling point?

(A) Ionic Bonds (B) London dispersion forces (C)Hydrogen Bonds (D) Metallic Bonds (E) Network Bonds

_____ 49. Hydrogen bonding would be seen in a sample of which of the following substances.

(A) CH_4 (B) H_2 (C) H_2O (D) HI (E) All of the above

_____ 50. Which of the following species does NOT have a tetrahedral structure?

(A) CH_4 (B) NH_4^+ (C) SF_4 (D) $AlCl_4^-$ (E) CBr_4

_____ 51. Which form of orbital hybridixation can form molecules with shapes that are either trigonal pyramidal or tetrahedral?

(A) *sp* (B) sp^2 (C) sp^3 (D) d^2sp (E) dsp^3

_____ 52. The size carbon atoms in a benzene molecule are shown in different resonance forms as three single bonds and three double bonds. If the length of a single carbon-carbon bond is 154 pm and the length of a double carbon-carbon bond is 133 pm, what length would be expected for the carbon-carbon bonds in benzene?

(A) 126 pm (B) 133 pm (C) 140 pm (D) 154 pm (E) 169 pm

FREE RESPONSE: Bonding and IMF's

Problem 53:

Explain each of the following in terms of atomic and molecular structures and/or intermolecular forces.

(a) Solid K conducts an electric current, whereas solid KNO_3 does not.

(b) $SbCl_3$ has measurable dipole moment, whereas $SbCl_5$ does not.

(c) The normal boiling point of CCl_4 is 77°C, whereas that of CBr_4 is 190°C.

(d) Iodine has a greater boiling point than bromine even though the bond energy in bromine is greater than the bond energy in iodine.

Problem 54:

Use appropriate chemical principles to account for each of the following observations. In each part, your response must include specific information about both substances.

(a) At 25°C and 1 atm, F_2 is a gas, whereas I_2 is a solid.

(b) The melting point of NaF is 993°C, whereas the melting point of CsCl is 645°C.

(c) The shape of the ICl_4^- ion is square planar, whereas the shape of the BF_4^- ion is tetrahedral

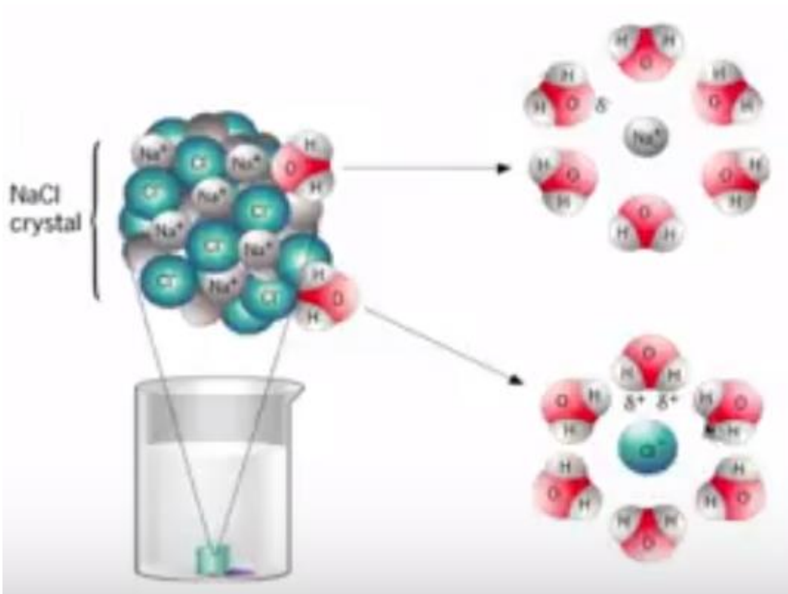
(d) Ammonia, NH_3, is very soluble in water, whereas phosphine, PH_3, is only moderately soluble in water.

Chromatography

More polar are more attracted to more polar substances

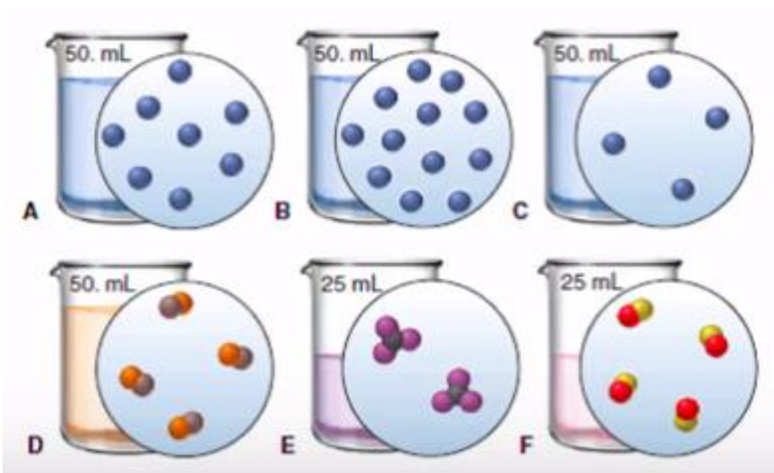
Dissolving/ Dissociation
Coulomb's Law helps determine solubility

Entropy in solutions



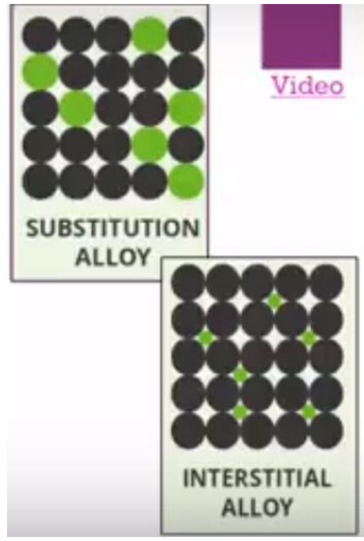
Molarity and Particle Views:

55. Rank the following in order of increasing molarity, each particle represents one mole. Some have equal concentrations.



Vapor Pressure

Distillation and Separating Solutions.

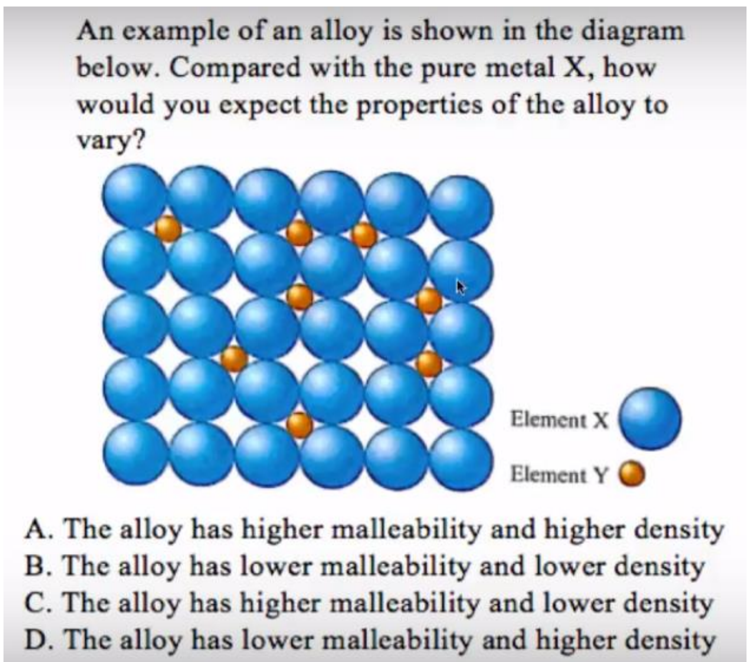


Alloys and their Properties

Substitutional
- Replacing one atom with a same sized atom

Interstitial
- Makes metals harder by plugging the holes by adding smaller atoms

______ 56.



Semiconductors

P type

S type

AP CHEMISTRY REVIEW: BIG IDEA # 3
CHEMICAL REACTIONS

Must Know:

Apply the law of conservation of mass to balance a chemical equation using symbols for atoms and molecules and particle diagrams.
Calculate the masses and moles of reactants and products using stoichiometry
Determine limiting reactants and percent yields from experimental data
Identify redox reactions and the electron transfer in redox reactions
Use an activity series to write balanced net ionic equations for redox reactions between metals and metal ions
Assign oxidation numbers to elements in a chemical formula
Perform limiting reactant and solution stoichiometry calculations
Interpret the results of a redox titration
Write balanced equations for the common reactions of Group 1 and 2 metals
Apply periodic properties to chemical reactivity
Identify REDOX reactions by the electrons transferred from the oxidized reactant to the reduced reactant
Assign an oxidation number to each atom in a chemical formula
Balance a redox reaction using the half-reaction method
Calculate quantities involved in a redox titration
Identify important redox reactions related to energy production such as the combustion of fossil fuels and the metabolism of food
Interpret diagram of voltaic and electrolytic cells
Identify oxidation at the anode and reduction at the cathode of an electrochemical cell
Calculate cell potential (voltage, EMF) under standard conditions using a table of standard reduction potentials
Compare qualitatively, using Le Chatelier's principle, the voltage and electron flow in a cell at nonstandard conditions to that of a cell at standard conditions
Predict spontaneity (thermodynamic favorability) from the cell potential of a redox reaction and its standard free energy, ΔG^o
Calculate quantities such as mass, current, time charge, and number of moles of electrons using Faraday's constant and the stoichiometry of a redox reaction.

Evidence of Chemical Change: production of gas, formation of a solid, changes in color, *production of heat

Combustion

Hydrocarbon + $O_2 \rightarrow CO_2 + H_2O$

Acid – Base Reactions

Neutralization (net ionic)

Practice: Write the net-ionic equation for the reaction between $KC_6H_7O_2$ + HCl

Bronsted – Lowery Acids and Bases

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$

Acid – donates a proton Base- accepts a proton
Amphoprotic

Precipitation Reactions

MUST form a solid from two solutions (net ionic)
Gravimetric analysis – precipitate something out to determine original concentration of ion
NAG SAG solubility Rules

Oxidation – Reduction Reactions

$2\,Mg + O_2 \rightarrow 2\,MgO$

Oxidation:
Reduction:

How many moles of electrons are transferred?

Zinc ions will react with aluminum metal according to the following chemical reaction:

$2\,Al + 2\,Zn^{+2} \rightarrow 2\,Al^{+3} + 3\,Zn$

Based on this chemical reaction, how many moles of electrons would be transferred when 1.0 mol of Zn^{+2} ions are consumed? (answer: 2 mol)

Redox Titrations: Use potassium permanganate because then you do not have to use an indicator because the Mn turns purple when it is no longer being reduced.

_____ 57. Galvanic/Voltaic Cell (E is + means spontaneous, occurring on it's own , Batteries have very high K and – G)

Half Cell 1: $Sn^{2+} + 2e^- \longrightarrow Sn$
Half Cell 2: $Ag^+ + e^- \longrightarrow Ag$
Half Cell 2: $Cr^{3+} + 3e^- \longrightarrow Cr$

Galvanic Cell	Half Cells	Reaction	$E°_{cell}$ (V)
X	1 & 2	$Sn + 2Ag^+ \longrightarrow 2Ag + Sn^{2+}$	0.94
Y	2 & 3	$Cr + 3Ag^+ \longrightarrow 3Ag + Cr^{3+}$	1.54
Z	1 & 3	$2Cr + 3Sn^{2+} \longrightarrow 3Sn + 2Cr^{3+}$	?

What is the cell potential of galvanic cell Z?

a. 0.26 V
b. 0.60 V
c. 2.48 V
d. 5.90 V

_____ 58. The following question is based on combining the three different half cells listed below:

Half Cell 1: $Sn^{2+} + 2e^- \longrightarrow Sn$
Half Cell 2: $Ag^+ + e^- \longrightarrow Ag$
Half Cell 2: $Cr^{3+} + 3e^- \longrightarrow Cr$

Galvanic Cell	Half Cells	Reaction	$E°_{cell}$ (V)
X	1 & 2	$Sn + 2Ag^+ \longrightarrow 2Ag + Sn^{2+}$	0.94
Y	2 & 3	$Cr + 3Ag^+ \longrightarrow 3Ag + Cr^{3+}$	1.54
Z	1 & 3	$2Cr + 3Sn^{2+} \longrightarrow 3Sn + 2Cr^{3+}$	?

In galvanic cells X and Z, which of the following takes place in half cell 1?

a. Oxidation occurs in both cell X and cell Z.
b. Reduction occurs in both cell X and cell Z.
c. Oxidation occurs in cell X and reduction in cell Z.
d. Reduction occurs in cell X and oxidation in cell Z.

_____ 59.

While cleaning up after the experiment, the student wishes to dispose of the unused solid I_2 in a responsible manner. The student decides to convert the solid I_2 to $I^-(aq)$ anion. The student has access to three solutions, $H_2O_2(aq)$, $Na_2S_2O_3(aq)$, and $Na_2S_4O_6(aq)$, and the standard reduction table shown below.

Half reaction	$E°$ (V)
$S_4O_6^{2-}(aq) + 2e^- \rightarrow 2S_2O_3^{2-}(aq)$	0.08
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	0.54
$O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O_2(aq)$	0.68

(e) Which solution should the student add to $I_2(s)$ to reduce it to $I^-(aq)$? Circle your answer below. Justify your answer, including a calculation of $E°$ for the overall reaction.

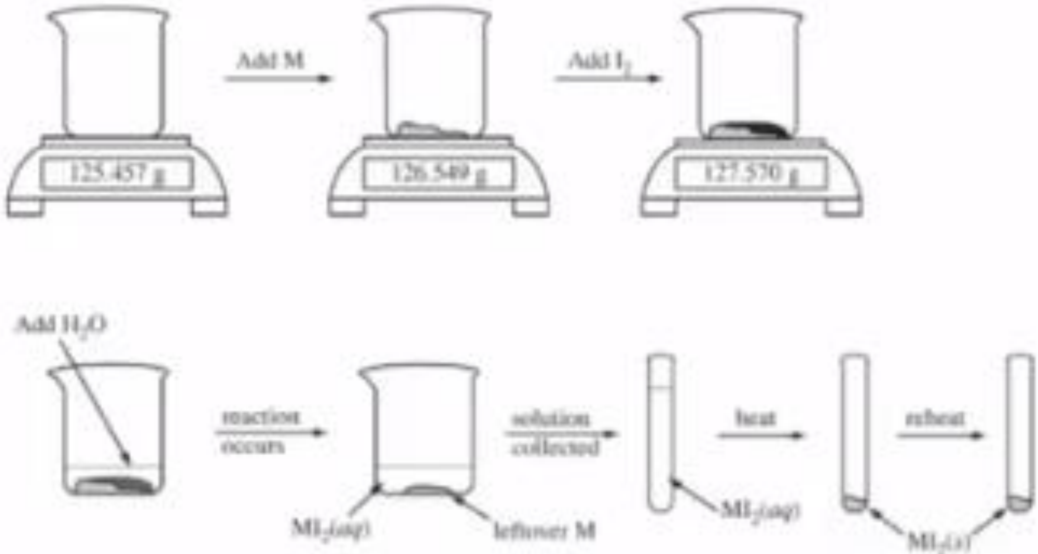
$H_2O_2(aq)$ $Na_2S_2O_3(aq)$ $Na_2S_4O_6(aq)$

60. Show calculation and so since E is positive it is thermodynamically favorable because E + makes G –

61. Write the balanced net ionic equation for the reaction between I_2 and chosen reactant.

Problem 62.

To determine the molar mass of an unknown metal, M, a student reacts iodine with an excess of the metal to form the water-soluble compound MI_2, as represented by the equation above. The reaction proceeds until all of the I_2 is consumed. The $MI_2(aq)$ solution is quantitatively collected and heated to remove the water, and th product is dried and weighed to constant mass. The experimental steps are represented below, followed by a data table.



Data for Unknown Metal Lab	
Mass of beaker	125.457 g
Mass of beaker + metal M	126.549 g
Mass of beaker + metal M + I_2	127.570 g
Mass of MI_2, first weighing	1.284 g
Mass of MI_2, second weighing	1.284 g

(a) Given that the metal M is in excess, calculate the number of moles of I_2 that reacted.

(b) Calculate the molar mass of the unknown metal M.

Making Predictions

Heat Carbonate always makes CO_2 gas.

$CaCO_3(s) \rightarrow CuO + CO_2$

Breaking Bonds is Endothermic

Percent Yield

Theoretical Yield is ALWAYS Stoichiometry

You can always improve procedure to heat longer to ensure all water is driven off.

Problem 63.

$Al_2S_3 + 6\ H_2O \rightarrow 2\ Al(OH)_3 + 3\ H_2S$

Using above equation, 15.00 g aluminum sulfide and 10.00 g water react.

(a)Identify the limiting reactant.

(b) What is the maximum mass of H_2S which can be formed from these reagents?

(c) How much excess reactant is left in container?

Experimental Design

Illustrate conservation of mass

Data Analysis

Always use mass of crucible after FINAL heating – to ensure all the water is gone.

Oxidation numbers must change because ELECTRONS MUST BE TRANSFERRED

_____ 64. $2\ N_2H_4(g) + N_2O_4(g) \rightarrow 3\ N_2(g) + 4\ H_2O(g)$

When 8.0 g of N_2H_4 (32 g mol^{-1}) and 92 g of N_2O_4 (92 g mol^{-1}) are mixed together and react according to the equation above, what is the maximum mass of H_2O that can be produced?

(a) 9.0 g (b) 18 g (c) 36 g (d) 72 g (e) 144 g

_____65.

$2\ H_2O(l) + 4\ MnO_4^{-}(aq) + 3\ ClO_2^{-}(aq) \rightarrow 4\ MnO_2(s) + 3\ ClO_4^{-}(aq) + 4\ OH^{-}(aq)$

According to the balance equation above, how many moles of $ClO_2^{-}(aq)$ are needed to react completely with 20. mL of 0.20 *M* $KMnO_4$ solution?

(a) 0.0030 mol (b) 0.0053 mol (c) 0.0075 mol (d) 0.013 mol (e) 0.030 mol

_____66. $3\ Ag(s) + 4\ HNO_3 \rightleftharpoons 3\ AgNO_3 + NO(g) + 2\ H_2O$

The reaction of silver metal and dilute nitric acid proceeds according to the equation above. If 0.10 mole of powdered silver is added to 10. milliliters of 6.0-molar nitric acid, the number of moles of NO gas that can be formed is

(a) 0.015 mole (b) 0.020 mole (c) 0.030 mole (d) 0.045 mole (e) 0.090 mole

_____67. A 20.0-milliliter sample of 0.200-molar K_2CO_3 solution is added to 30.0 milliliters of 0.400-molar $Ba(NO_3)_2$ solution. Barium carbonate precipitates. The concentration of barium ion, Ba^{2+}, in solution **after** reaction is

(a) 0.150 M (b) 0.160 M (c) 0.200 M (d) 0.240 M (e) 0.267 M

_____68. What number of moles of O_2 is needed to produce 142 grams of P_4O_{10} from P? (Molecular weight P_4O_{10} = 284)

(a) 0.500 mole (b) 0.625 mole (c) 1.25 mole (d) 2.50 mole (e) 5.00 mole

_____69. $BrO_3^- + 5\ Br^- + 6\ H^+ \rightleftharpoons 3\ Br_2 + 3\ H_2O$

If 25.0 milliliters of 0.200-molar BrO_3^- is mixed with 30.0 milliliters of 0.450-molar Br^- solution that contains a large excess of H^+, the amount of Br_2 formed, according to the equation above, is

(a) 5.00×10^{-3} mol (b) 8.10×10^{-3} mol (c) 1.35×10^{-2} mol (d) 1.50×10^{-2} mol (e) 1.62×10^{-2} mol

____70. Commercial vinegar was titrated with NaOH solution to determine the content of acetic acid, $HC_2H_3O_2$. For 20.0 milliliters of the vinegar 26.7 milliliters of 0.600-molar NaOH solution was required. What was the concentration of acetic acid in the vinegar if no other acid was present?

(a) 1.60 M (b) 0.800 M (c) 0.600 M (d) 0.450 M (e) 0.200 M

_____71. What volume of 0.150-molar HCl is required to neutralize 25.0 millilters of 0.120-molar $Ba(OH)_2$?

(a) 20.0 mL (b) 30 0 mL (c) 40.0 mL (d) 60.0 mL (e) 80.0 mL

_____72. It is suggested that SO_2 (molar mass 64 grams), which contributes to acid rain, could be removed from a stream of waste gases by bubbling the gases through 0.25-molar KOH, thereby producing K_2SO_3. What is the maximum mass of SO_2 that could be removed by 1,000. liters of the KOH solution?

(a) 4.0 kg (b) 8.0 kg (c) 16 kg (d) 20. Kg (e) 40. kg

_____73. Commercial vinegar was titrated with NaOH solution to determine the content of acetic acid, $HC_2H_3O_2$. For 20.0 milliliters of the vinegar, 32.0 milliliters of 0.500-molar NaOH solution was required. What was the concentration of acetic acid in the vinegar if no other acid was present?

(a) 1.60 M (b) 0.800 M (c) 0.640 M (d) 0.600 M (e) 0.400 M

_____74. What is the final concentration of barium ions, $[Ba^{2+}]$, in solution when 100. mL of 0.10 M $BaCl_2(aq)$ is mixed with 100. mL of 0.050 M $H_2SO_4(aq)$?

(a) 0.00 M (b) 0.012 M (c) 0.025 M (d) 0.075 M (e) 0.10 M

MULTIPLE CHOICE REVIEW: General Mole Relationships

_____75. The atomic mass of copper is 63.55. Given that there are only two naturally occurring isotopes of copper, ^{63}Cu and 65Cu, the natural abundance of the 65Cu isotope must be approximately

(a) 90% (b) 70% (c) 50% (d) 25% (e) 10%

_____76. What is the mole fraction of ethanol, C_2H_5OH, in an aqueous solution that is 46 percent ethanol by mass? The molar mass of C_2H_5OH is 46 g, the molar mass of H_2O is 18 g.)

(a) 0.25 (b) 0.46 (c) 0.54 (d) 0.67 (e) 0.75

_____77. Approximately what mass of $CuSO_4 \bullet 5\ H_2O$ (250 g mol^{-1}) is required to prepare 250 mL of 0.10 M copper(II) sulfate solution?

(a) 4.0 g (b) 6.2 g (c) 34 g (d) 85 g (e) 140 g

_____ 78. If 200. mL of 0.60 *M* $MgCl_2(aq)$ is added to 400. mL of distilled water, what is the concentration of $Mg^{2+}(aq)$ in the resulting solution? (Assume volumes are additive.)

(a) 0.20 *M* (b) 0.30 *M* (c) 0.40 *M* (d) 0.60 *M* (e) 1.2 *M*

_____79. The weight of H_2SO_4 (molecular weight 98.1) in 50.0 milliliters of a 6.00-molar solution is

(a) 3.10 grams (b) 12.0 grams (c) 29.4 grams (d) 294 grams (e) 300. grams

_____80. How many grams of calcium nitrate, $Ca(NO_3)_2$, contains 24 grams of oxygen atoms?

(a) 41 grams (b) 50.grams (c) 62 grams (d) 96 grams (e) 164grams

_____81. The mass of element Q found in 1.00 mole of each of four different compounds is 38.0 grams, 57.0 grams, 76.0 grams, and 114 grams, respectively. A possible atomic weight of Q is

(a) 12.7 (b) 19.0 (c) 27.5 (d) 38.0 (e) 57.0

_____82. When 70. milliliter of 3.0-molar Na_2CO_3 is added to 30. milliliters of 1.0-molar $NaHCO_3$ the resulting concentration of Na^+ is

(a) 2.0 M (b) 2.4 M (c) 4.0 M (d) 4.5 M (e) 7.0 M

_____83. Mass of an empty container = 3.0 grams
Mass of the container plus the solid sample = 25.0 grams
Volume of the solid sample = 11.0 cubic centimeters

The data above were gathered in order to determine the density of an unknown solid.
The density of the sample should be reported as

(a) 0.5 g/cm^3 (b) 0.50 g/cm^3 (c) 2.0 g/cm^3 (d) 2.00 g/cm^3 (e) 2.27 g/cm^3

_____ 84. If 87 grams of K_2SO_4 (molar mass 174 grams) is dissolved in enough water to make 250 milliliters of solution, what are the concentrations of the potassium and the sulfate ions?

	$[K^+]$	$[SO_4^{2-}]$
(A)	0.020 M	0.020 M
(B)	1.0 M	2.0 M
(C)	2.0 M	1.0 M
(D)	2.0 M	2.0 M
(E)	4.0 M	2.0 M

_____ 85. When a 1.25-gram sample of limestone was dissolved in acid, 0.44 gram of CO_2 was generated. If the rock contained no carbonate other than $CaCO_3$, what was the percent of $CaCO_3$ by mass in the limestone?

(a) 35% (b) 44% (c) 67% (d) 80% (e) 100%

_____86. What mass of Au is produced when 0.0500 mol of Au_2S_3 is reduced completely with excess H_2?

(a) 9.85 g (b) 19.7 g (c) 24.5 g (d) 39.4 g (e) 48.9 g

_____87. A 1.0 L sample of an aqueous solution contains 0.10 mol of NaCl and 0.10 mol of $CaCl_2$. What is the minimum number of moles of $AgNO_3$ that must be added to the solution in order to precipitate all of the Cl^- as AgCl(s)? (Assume that AgCl is insoluble.)

(a) 0.10 mol (b) 0.20 mol (c) 0.30 mol (d) 0.40 mol (e) 0.60 mol

_____88. The volume of distilled water that should be **added** to 10.0 mL of 6.00 M HCl (aq) in order to prepare a 0.500 M HCl (aq) solution is approximately

(a) 50.0 mL (b) 60.0 mL (c) 100. mL (d) 110. mL (e) 120. mL

_____89. How many milliliters of 11.6-molar HCl must be diluted to obtain 1.0 liter of 3.0-molar HCl?

(a) 3.9 mL (b) 35 mL (c) 250 mL (d) 1,000 mL (e) 3,900 mL

_____90. A measured mass of an unreactive metal was dropped into a small graduated cylinder half filled with water. The following measurements were made.

Mass of metal = 19.611 grams
Volume of water before addition of metal = 12.4 milliliters
Volume of water after addition of metal = 14.9 milliliters

The density of the metal should be reported as

(a) 7.8444 g/mL (b) 7.844 g/mL (c) 7.84 g/mL (d) 7.8 g/mL (e) 8 g/mL

_____91. How many moles of solid $Ba(NO_3)_2$ should be added to 300. milliliters of 0.20-molar $Fe(NO_3)_3$ to increase the concentration of the NO_3^- ion to 1.0-molar? (Assume that the volume of the solution remains constant.)

(a) 0.060 mole (b) 0.12 mole (c) 0.24 mole (d) 0.30 mole (e) 0.40 mole

_____92. In which of the following compounds is the mass ratio of chromium to oxygen closest to 1.6 to 1.0?

(a) CrO_3 (b) CrO_2 (c) CrO (d) Cr_20 (e) Cr_2O_3

AP CHEMISTRY REVIEW: Electrochemistry

Redox Reaction is a reaction where there is a change in oxidation numbers. Some atoms will show an increase and some will show a decrease in oxidation number.
Oxidation Numbers are assigned according to the rules studied earlier.

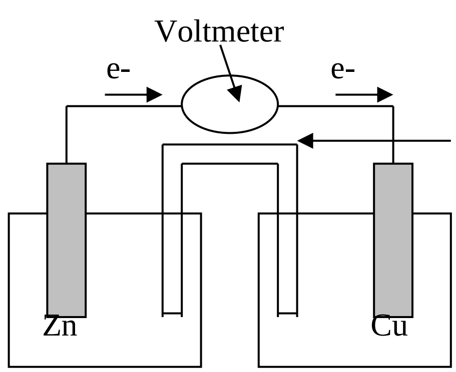
1. The oxidation number of an element in the free, or un-combined, state is *zero.*
2. The oxidation number of a *monatomic ion* is equal to its charge.
3. In all compounds containing *Group IA alkali metals*, the oxidation number of the Group IA ion is +1.
4. In all compounds containing *Group IIA metals,* the oxidation number of the Group IIA ion is +2.
5. In most compounds containing *oxygen*, the oxidation number of oxygen is almost always -2.
6. In most compounds containing *hydrogen*, the oxidation number of hydrogen is almost always +1.
7. The algebraic sum of the oxidation numbers of all the atoms in the formula of a compound is zero.
8. The algebraic sum of the oxidation numbers of all the atoms in the formula of a polyatomic ion is equal to the charge on the ion

Oxidation is the loss of electron(s). The oxidation number increases.
Half rxn is: $M \rightarrow M^+ + \mathbf{e^-}$

Reduction is the gaining of electron(s). The oxidation number decreases.
Half rxn is: $M^+ + \mathbf{e^-} \rightarrow M$

Helpful Phrase to remember:
A LEO goes **C GER** (**A**node: **L**oss of **E**lectron is **O**xidation), (**C**athode: **G**ain of **E**lectron is **R**eduction)

Galvanic, Electrochemical, or Voltaic Cells



Salt Bridge (KNO_3, saturated)
=maintain charge neutrality in half cells

$Zn_{(s)} \rightarrow Zn^{+2} + 2e^-$ — $Cu^{+2} + 2e^- \rightarrow Cu_{(s)}$
Oxidation Half-Cell — Reduction Half-Cell
Anode — Cathode
$E^o = +.76$ volts — $E^o = +.34$ volts
$\underline{E^o_{cell} = +1.10 \text{ volts}}$

(Line Notation) $Zn_{(s)}\,|Zn^{2+}(aq,\ 1M)|\ KNO_3(sat.)|Cu^{2+}(aq,1M)|Cu_{(s)}$

Properties Galvanic Cell:

- The flow of electrons provides a voltage potential difference between the two half-cells.
- This is the **Electromotive force** or **emf** ($E°_{cell}$). A positive voltage means a galvanic cell will operate spontaneously.
- **Standard Reduction Potential** (E^o)= Solutions are at 1*Molar*, gases are at 1 atm and Temperature is 25 °C.
- When comparing **two half reduction** equations, the one with the Largest $E°_{red}$ is Reduced. The other one is Oxidized, which then must be reversed and the sign of $E°_{red}$ reversed to represent $E°_{oxidation}$.
- Cell Potential: $E^o_{cell} = E°_{reduction} + E°_{oxidation}$
- Changing the stoichiometric coefficients of a half-cell *does not change* the value of E^o.

Anode Half Cell Description:

1. oxidation is occurring resulting in the production of positive ions and a build up of positive charges in the solution.
2. The metal anode electrode is being oxidized so it will lose mass.

Cathode Half Cell Description:

1. Reduction is occurring there, so the Metal Cations from the solution are converted to Metal(s). This results in a lost of positive charges and a buildup of negative charges in the solution.
2. The newly reduced metal(s) stick to the cathode resulting in increased mass.

Salt Bridge: It must contain a saturated soluble ionic solution.

1. The salt bridge connects the two half cells and allows for IONS to flow between them. The ion flow helps maintain CHARGE neutrality in each half cell.
2. Anions flow to the Anode side, and Cations flow to the Cathode side.

Wire that connects the Anode and the Cathode:

1. Electrons ALWAYS flow from the Anode to the Cathode through the wire.

Free Energy, Cell Potential and Equilibrium

$\Delta G^o = -nFE^o_{cell}$ (1 Faraday(F) = 96,500 C/mol, *n = # mole e- transferred*)

$\Delta G^o = -RT\ln K$ (R = 8.314 J/Kmol)

(Note: When E_{cell} is positive, the reaction is spontaneous and ΔG^o is negative)

Effect of Change in Concentration on E_{cell}: Explain using Le Chatalier's Principle:

Electrolysis: (Electrolytic Cell) is nonspontaneous redox reactions. Ecell < 0

1. Identify the reactants present in the electrolytic cell.
2. Write half-reactions for each reactant present.
3. (aqueous) vs. molten

Calculations

Electrical energy (ampere = C/s) is used to cause a non-spontaneous reaction to occur.

General Calculations:

$$amp\left(\frac{C}{s}\right) \text{x } time(s) \text{ x } \frac{1 mole\ e^-}{96{,}500C} \text{ x } \frac{\text{1 mole Element}}{\text{\# mole e}^-} = \# mole \text{ Element}$$

Coulombs(C) is Charge

1 Faraday (F) = 96,500 C = 1mole e-

Problem 93. A current of .452 amps is passed through an electrolytic cell containing molten $CaCl_2$ for 1.5 hours. (a) Write the electrode reactions and

(b) calculate the quantity of products (in grams) formed at the respective electrodes.

(c) If this reaction was aqueous as opposed to molten would the reactions be the same or different, how come?

MULITPLE CHOICE REVIEW: Electrochemistry

_____94. In the electroplating of nickel, 0.200 faraday of electrical charge is passed through a solution of $NiSO_4$. What mass of nickel is deposited?

(a) 2.94 g (b) 5.87 g (c) 11.7 g (d) 58.7 g (e) 294 g

_____95. If 0.060 faraday is passed through an electrolytic cell containing a solution of In^{3+} ions, the maximum number of moles of In that could be deposited at the cathode is

(a) 0.010 mole (b) 0.020 mole (c) 0.030 mole (d) 0.060 mole (e) 0.18 mole

____96. When the below equation is correctly balanced the sum of the reactant's coefficients is

$$\_\_\_H^+ + \_\_\_MnO_4^- + \_\_\_C_2O_4^{2-} \rightarrow \_\_\_H_2O + \_\_\_CO_2 + \_\_\_Mn^{2+}$$

(A) 8 (B) 10 (C) 14 (D) 23 (E) 38

$$Zn(s) + Cu^{+2}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

____97. The equilibrium constant, *K*, for the reaction above is greater than 1×10^{10}. Which of the following correctly describes the standard voltage, $E°$, and the standard free energy change, $\Delta G°$, for this reaction?

(A) $E°$ is positive and $\Delta G°$ is negative.
(B) $E°$ is negative and $\Delta G°$ is positive.
(C) $E°$ and $\Delta G°$ are both positive.
(D) $E°$ and $\Delta G°$ are both negative.
(E) $E°$ and $\Delta G°$ are both zero.

Questions 95 and 96 refer to an electrolytic cell with the following half reaction which occurs in an aqueous solution of $Cr(NO_3)_3$.

$$Cr^{3+} + 3\ e^- \rightarrow Cr(s)$$

______98. A current is passed though a $Cr(NO_3)_3(aq)$ solution for 1 hour and 15 minutes. After this time the mass of chromium metal produced was 8.083 grams. Which of the following expressions correctly shows how to calculate the amount of current, in amperes, required? (1 faraday = 96,500 coulombs)

(A) $\dfrac{(8.083)(3)(96500)}{(75)(60)(52.00)}$

(B) $\dfrac{(96500)(52.00)}{(8.083)(75)(60)(3)}$

(C) $\dfrac{(8.083)(75)(52.00)}{(3)(96500)}$

(D) $\dfrac{(52.00)(3)}{(96500)(8.083)(75)(60)}$

(E) $\dfrac{(8.083)(75)(60)(52.00)}{(3)(96500)}$

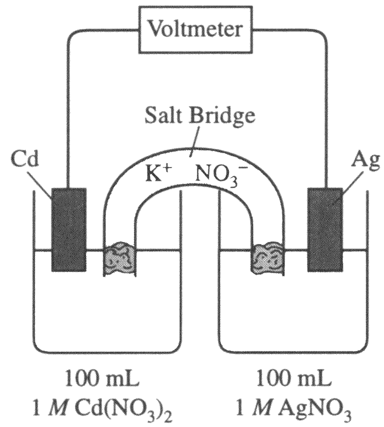
_______99. Which of the following occurs in this half reaction?

I. The reaction is spontaneous.
II. Cr^{3+} is reduced at the cathode.
III. Water is oxidized at the anode

(A) I only (B) II only (C) III only (D) I and III only (E) II and III only

FREE RESPONSE REVIEW: Electrochemistry

Problem 100:



Answer the following questions regarding the electrochemical cell shown.

(a) Write the balanced net-ionic equation for the spontaneous reaction that occurs as the cell operates, and determine the cell voltage.

(b) In which direction do anions flow in the salt bridge as the cell operates? Justify your answer.

(c) If 10.0 mL of 3.0-molar $AgNO_3$ solution is added to the half-cell on the right, what will happen to the cell voltage? Explain.

(d) If 1.0 gram of solid NaCl is added to each half-cell, what will happen to the cell voltage? Explain.

(e) If 20.0 mL of distilled water is added to both half-cells, the cell voltage decreases. Explain.

Problem 101:

$$Sr(s) + Mg^{2+} \rightarrow Sr^{2+} + Mg(s)$$

Consider the reaction represented above that occurs at 25°C. All reactants and products are in their standard states. The value of the equilibrium constant, K_{eq}, for the reaction is 4.2×10^{17} at 25°C.

(a) Predict the sign of the standard cell potential, $E°$, for a cell based on the reaction. Explain your prediction.

(b) If the reaction were carried out at 60°C instead of 25°C, how would the cell potential change? Justify your answer in terms of LeChatelier's Principle.

(c) How would the cell potential change if the reaction were carried out at 25°C with a 1.0-molar solution of $Mg(NO_3)_2$ and a 0.10-molar solution of $Sr(NO_3)_2$? Explain in terms of LeChatelier's Principle.

(d) When the cell reaction in (d) reaches equilibrium, what is the cell potential?

Problem 102:

In an electrolytic cell, a current of 0.250 ampere is passed through a solution of a chloride of iron, producing Fe*(s)* and Cl_2*(g)*.

(a) Write the equation for the half-reaction that occurs at the anode.

(b) When the cell operates for 2.00 hours, 0.521 gram of iron is deposited at one electrode. Determine the formula of the chloride of iron in the original solution.

(c) Write the balanced equation for the overall reaction that occurs in the cell.

(d) How many liters of Cl_2*(g)*, at 25°C and 750 mm Hg, are produced when the cell operates as described in part (b)?

(e) Calculate the current that would produce chlorine gas from the solution at a rate of 3.00 grams per hour.

AP CHEMISTRY REVIEW: BIG IDEA # 4
Kinetics

Must Know:

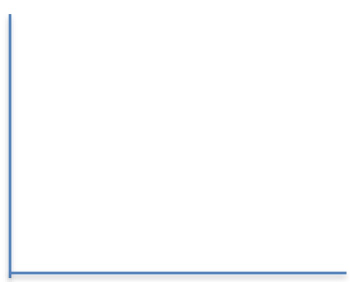
- Define rate of reaction and list common ways to express the rate of reaction
- Design an experiment or interpret experimental data that measure rate
- Explain the factors that affect the rate of reaction
- Use the collision model to justify how concentration, pressure, temperature, and the phase of the reactants affect reaction rates.
- Examine concentration versus reaction rate data using the method of initial rates to determine the rate law and the order of reaction
- Describe a rate constant and explain how it characterizes a reaction
- Connect the half-life of first order reaction to its rate constant
- Infer reaction order from plots of concentration vs. time data
- Justify how temperature affects rate constant
- Interpret Maxwell-Boltzmann plots that describe distributions of particle energies
- Use energy profiles to make qualitative predictions about the relative rates of reaction
- Evaluate reaction mechanisms and determine which are consistent with rate data
- Interpret data that infer the presence of reaction intermediates
- Explain how catalyst work, including acid-base catalysts, surface catalysts and enzymes

Collision Theory: In order for a reaction to occur particles must collide with sufficient kinetic energy and proper orientations.

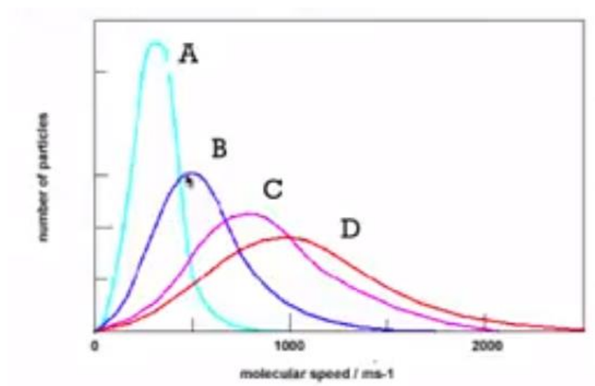
Ways to Increase # of collisions:
- Surface Area
- Concentration
- Ionic are faster than molecular
- Pressure (gas only)
- Temperature

Catalyst:

Rate =



Time (s)



Relative Rates: Relate Rates to each other

$$2\,A + 3\,B \rightarrow C + 2\,D$$

a.) Differential Rate: Rate versus Concentrations

Solve using: Table Logic

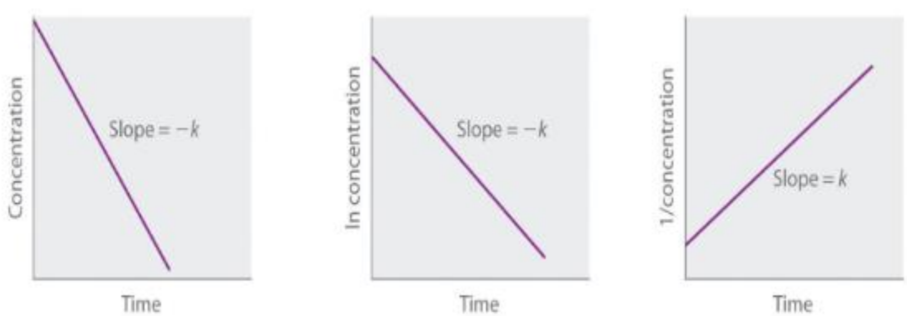
Experiment	[A]	[B]	Rate of disappearance of [A] M/time
1	.01	.02	5
2	.01	.04	10
3	.02	.04	20

Determine the "Rate Law" or "Rate Expression"

e.) Integrated Rate Law: Rate versus Time

y = mx + b

0 order **1^{st} order** **2^{nd} order**



CHEM-IS-TRY, WORK. 38

MULTIPLE CHOICE REVIEW: Kinetics

Rate = $k[M][N]^2$

_____103. The rate of a certain chemical reaction between substances M and N obeys the rate law above. The reaction is first studied with [M] and [N] each 1×10^{-3} molar. If a new experiment is conducted with [M] and [N] each 2×10^{-3} molar, the reaction rate will increase by a factor of

(a) 2 (b) 4 (c) 6 (d) 8 (e) 16

_____ 104. The rate law for a reaction is found to be Rate = $k[A]^2[B]$. What is the intermediate? Which of the following mechanisms gives this rate law?

I. $A + B \rightleftarrows E$ (fast)
$E + B \rightarrow C + D$ (slow)

II. $A + B \rightleftarrows E$ (fast)
$E + A \rightarrow C + D$ (slow)

III. $A + A \rightarrow E$ (slow)
$E + B \rightarrow C + D$ (fast)

A. I
B. II
C. III
D. Two of these

105: Draw and label axes for the energy profiles below. Match the curves with the appropriate description.

A. **exothermic reaction with a 2 step mechanism where the first step is slow.**

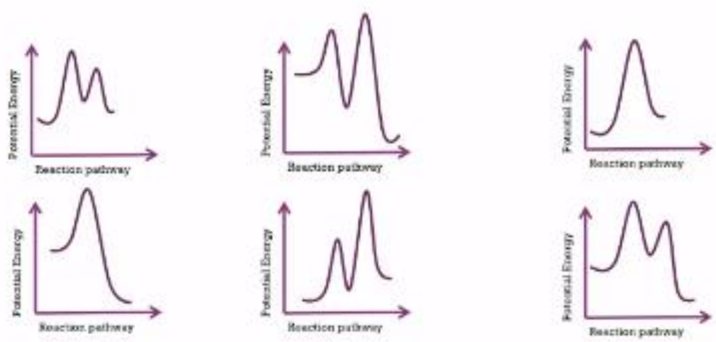
B. **endothermic reaction with a 2 step mechanism where the second step is slow**

C. **exothermic reaction with a 2 step mechanism where the second step is slow.**

D. **endothermic reaction with a 2 step mechanism where the first step is slow.**

E. **exothermic reaction with a 1 step mechanism.**

F. **endothermic reaction with a 1 step mechanism.**



FREE RESPONSE REVIEW: Kinetics

Problem 106: Use the experimental data provided below to answer the following set of questions.

$$A + B \rightarrow 2C + D$$

Exp.	[A]	[B]	Initial rate disappearance (M/s)
1	0.033	0.034	6.67×10^{-2}
2	0.034	0.137	1.08×10^{-2}
3	0.136	0.136	1.07×10^{-2}
4	0.202	0.233	?

1. Write the order of the reaction with respect to
   a. A

   b. B

2. Write the rate expression.

3. Solve for k, including units.

4. Determine the rate in the fourth experiment

5. Determine the rate of formation of change of [C] in experiment 3.

6. Elementary Step 1 $B + B \rightarrow E + D$ slow
   Elementary Step 2 $E + A \rightarrow B + C$ fast, equilibrium

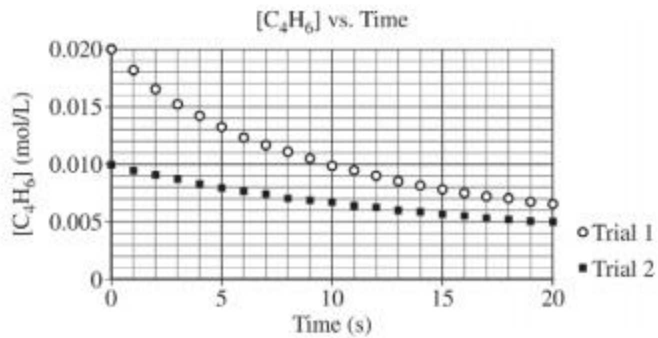
   a. Is the above proposed mechanism plausible? Explain.

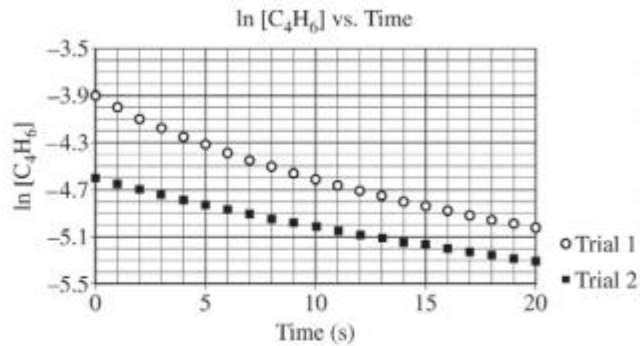
   b. Is E an intermediate or a catalyst, explain.

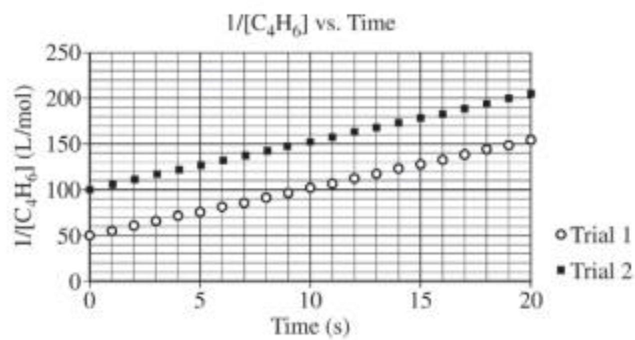
Problem 107:

$$2\ C_4H_6(g) \rightarrow C_8H_{12}(g)$$

At high temperatures the compound C_4H_6 (1,3-butadiene) reacts according to the equation above. The rate of the reaction was studied at 625 K in a rigid reaction vessel. Two different trials, each with a different starting concentration, were carried out. The data were plotted in three different ways, as shown below.







(a) For trial 1, calculate the initial pressure, in atm, in the vessel at 625 K. Assume that initially all the gas present in the vessel is C_4H_6.

(b) Use the data plotted in the graphs to determine the order of the reaction with respect to C_4H_6.

(c) The initial rate of the reaction in trial 1 is 0.0010 mol/(L·s). Calculate the rate constant, k, for the reaction at 625 K.

AP CHEMISTRY REVIEW: BIG IDEA # 5
Thermochemistry

Must Know:

Use molecular collisions to explain or predict the transfer of heat between systems
Use the law of conservation of energy to explain or predict the transfer of heat between systems
Explain the quantity of energy change that occurs when two substances of different temperatures interact
Calculate or estimate enthalpy changes associated with chemical reactions
Calculate or estimate energy changes associated with temperature changes using heat capacity
Use results of a constant pressure calorimetry experiment to determine the change in enthalpy of a chemical or physical process
Use bond energies to calculate or estimate enthalpies of reaction
Distinguish between exothermic and endothermic changes and understand the direction of energy flow for each
Explain using the kinetic-molecular theory the concept of entropy in a chemical system
Explain the direction and relative magnitudes of changes in entropy when phase changes occur in solids, liquids or gases
Use models and representations to explain how molecular complexity affects entropy
Calculate entropy changes from thermodynamic data
Use $\Delta G^o = \Delta H^o - T\Delta S^o$ to predict thermodynamic favorability of a chemical or physical change
Calculate G to determine the thermodynamic favorability of a chemical change
Use the relationship between G and K to estimate the magnitude of K and the thermodynamic favorability process
Explain how the technique of chromatography uses intermolecular attractions to separate mixtures

AB: Bond Energy is ALWAYS Endothermic

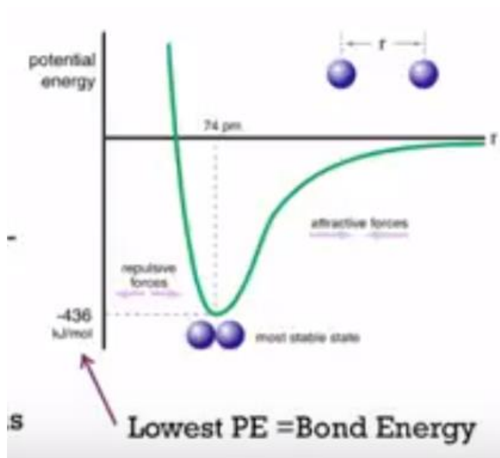
3 Factors:

Size

Polarity

Bond Order

Radius providing most stable energy for a bond

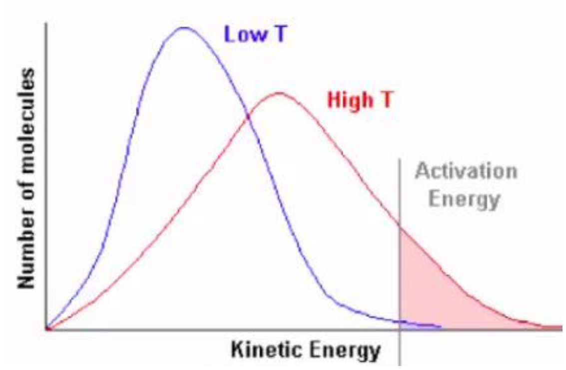


108. The bond energy of the carbon to sulfur bond in CS_2 is 577kJ/mol. Is the bond energy of the carbon selenium bond in CSe_2 expected to be greater than, less than or equal to its value. Justify your answer.

109. Consider the three gases in the take at 327°C: CH_3OH, CO and H_2

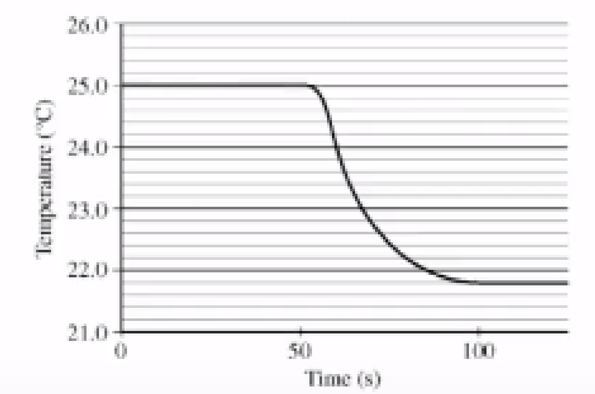
110) How do the average kinetic energies of the molecules of the gases compare? Explain.

111) Which gas has the highest average molecular speed? Explain.



Universe: The sum of the system and surroundings
System: The species we are studying
Surroundings: The environment outside of the system (generally, we take temperature of this...)
Endothermic: Heat flows to the system from surroundings
Exothermic: Heat flows from the system to the surroundings

A student performs an experiment to measure the molar enthalpy of a solution of urea, H_2NCONH_2. The student places 91.05 g of water at 25°C. Into a coffee cup calorimeter and immerses a thermometer in the water. After 50 s the student adds 5.13 g of solid urea, also at 25°C, to the water and measures the temperature of the solution as the urea dissolves. A plot of the temperature data is shown in the graph below.



112. According to the data is the dissolution of urea in water an endothermic process or an exothermic process. Justify your answer.

Heat Transfer: Energy is always transferred from high to low, when temperatures are equal they are at thermo equilibrium.

Specific Heat Capacity: amount of heat required to heat up 1 gram of a substance by 1 °C.

113. Assume that the specific heat capacity of the calorimeter is negligible and that the specific heat capacity of the solution of urea and water is 4.3 J/g°C throughout the experiment.
Calculate the heat of dissolution of the urea in joules.

1st Law: Energy is conserved.
$Q_{lost} + Q_{gained} = O$

A student investigates the enthalpy of solution for two alkali metal halides, LiCl and NaCl. In addition to the alts, the student has access to a calorimeter, a balance with a precision of 0.1 g and a thermometer with a precision of 0.1ºC.

114. To measure heat of solution for LiCl the student adds 100.0 g of water initially at 15.0ºC to a calorimeter and adds 10.0 g of LiCl(s), stirring to dissolve. After the LiCl dissolves completely, the maximum temperature reached by the solution is 35.6ºC.

a. Calculate the magnitude of the heat absorbed by the solution during the dissolution process, assuming that the specific heat capacity of the solution is 4.18 J/gºC. Include units with your answer.

Heating Curve

Cooling Curve

Electrostatic forces exist between molecules as well as between atoms or ions, and breaking these intermolecular interactions requires energy.

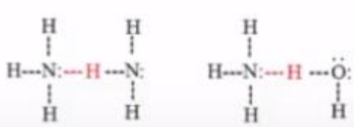
The ***Stronger the IMF*** the more energy required to break it, the ***Higher the Boiling Point,*** the ***Lower the Vapor Pressure***.

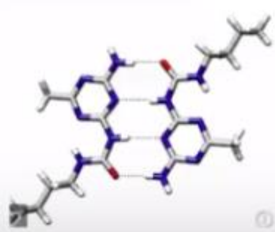
Intermolecular Forces Listed from weakest to strongest. Thus the boiling points and vapor pressure of molecular substances can be ordered based on IMF strength:

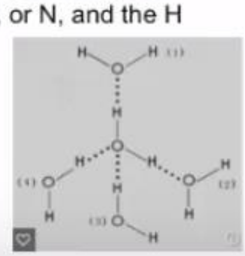
Weaker IMF, Lower Boiling, Higher Vapor Pressure

1. **Dispersion (Induced Dipole- Induced Dipole):** Caused by distortion of electron cloud. The larger the electron cloud, and the more surface area, the more polarizable the cloud, the stronger the dispersion force. ***Thus the boiling point trend in halogens is $I_2 > Br_2 > Cl_2 > F_2$ and n-butane (30.2° C) has a higher boiling point than isobutane (-11 °C). All substances have dispersion forces, as all electron clouds distort.*** Nonpolar molecules and atoms have only dispersion forces, as they have no permanent dipoles.
2. **Dipole- Induced Dipole**: Occurs between a polar molecule (HCl) and a nonpolar molecule. (Cl_2) The nonpolar molecule's cloud distorts when affected by a dipole.
3. **Dipole-Dipole:** Occurs between 2 polar molecules. (HCl-HCl)
4. **Hydrogen Bond**: An extreme case of Dipole – Dipole. Occurs between molecules containing a H covalently bonded to F,O, or N. The ***"bond"*** occurs between the lone pair of F, O, or N, and the H which is attached to one of those elements.

Stronger IMF, Higher Boiling, Lower Vapor Pressure







115. Why does dimethyl ether boil at a lower boiling point than ethanol

116. Rank molecules based on boiling points.

Physical vs. Chemical Changes

Coupling of reactions; When we couple reactions we may multiply by a factor or reverse the reaction before we couple or add them together.

Manipulation	Impact on Equilibrium Constant, K	Impact on Free Energy (G) , Enthalpy (H) or Entropy (S)	Impact on Standard Reduction Potential (E)
Multiply by factor	Raise K to power of factor	Multiply by factor	NOTHING
Flip/Reverse Reaction	Inverse of K aka 1/K	Change Signs	Change sign
Add reactions	Multiply values	Add values	Add values

AP CHEMISTRY REVIEW: Thermochemistry

1st Law- energy of the universe is constant
2nd law- if process is spontaneous in one direction it cannot be spontaneous in the reverse direction.

Calorimeter (Temp container constant)

In a regular calorimeter assume the container is not absorbing any heat.

$$0 = Q_{\text{reaction}} + Q_{\text{water}}$$

$$\mathbf{Q}_{\text{reaction}} = -(\mathbf{m}_{H_2O}\mathbf{C}_{H_2O}\Delta\mathbf{T}_{H_2O})$$

State Functions: depend ONLY on change between initial and final states, not on process by which change occurs (pathway does not matter).

1.) Enthalpy, ΔH

Enthalpy change:

A measure of the energy that is released or absorbed by the substance when bonds are broken and formed during a reaction.

Heat of formation, ΔH_f

Change in Energy that takes place when one mole of a compound is formed from its component pure elements under standard state conditions.

ΔH_f for an **element** is always **zero!**
ΔH_f for a compound (-) means exothermic, product lower E, more stable
ΔH_f for a compound (+) means endothermic, product higher E, less stable

$\Delta H^o = \sum \Delta H^o_{product} - \sum \Delta H^o_{reactants}$

Hess's Law: ΔH for overall reaction is simply the sum of the ΔH values for all steps.

- If you reverse reaction, change the sign on ΔH
- If you divide reaction by 2, divide ΔH by 2.

2.) Entropy, ΔS

Measure of randomness or disorder
ALL substances have a $+\Delta S$ value because 0 entropy is defined as a solid crystal at 0K.

gas > solution > pure liquid > solid

$\Delta S^o = \sum \Delta S^o_{product} - \sum \Delta S^o_{reactants}$

3.) Free Energy, ΔG

Measures free energy and predicts the spontaneity of the reaction.
ΔG (-) reaction is spontaneous
ΔG (+) reaction is NOT spontaneous
$\Delta G = 0$, reaction is at equilibrium.

$\Delta G^o = \sum \Delta G^o_{product} - \sum \Delta G^o_{reactants}$

$\Delta G^o \quad = \quad \Delta H^o \quad - \quad T\,\Delta S^o$

$\Delta G = \quad \Delta G^o + RT \ln Q$

at Equilibrium $\Delta G = 0$, $Q = K$

MULTIPLE CHOICE REVIEW: Thermochemistry

_____117. $3\ C_2H_2(g) \rightarrow C_6H_6$

What is the standard enthalpy change, ΔH^o, for the reaction represented above?

ΔH_f of $C_2H_2(g)$ is 230 kJ mol^{-1} ; ΔH_f of $C_6H_6(g)$ is 83 kJ mol^{-1}

(a) -607 kJ (b) -147 kJ (c) -19 kJ (d) +19 kJ (e) +773 kJ

_____118. $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$; $\Delta H = -\ 889.1$ kJ

$\Delta H_f°\ H_2O(l) = -\ 285.8$ kJ / mole $\Delta H_f°\ CO_2(g) = -\ 393.3$ kJ / mole

What is the standard heat of formation of methane, $\Delta H_f°\ CH_4(g)$, as calculated from the data above?

(a) -210.0 kJ/mole (b) -107.5 kJ/mole (c) -75.8 kJ/mole (d) 75.8 kJ/mole (e) 210.0 kJ/mole

_____119. $I_2(g) + 3\ Cl_2(g) \rightarrow 2\ ICl_3(g)$

According to the data in the table below, what is the value of $\Delta H°$ for the reaction represented above?

Bond	Average Bond Energy (kilojoules / mole)
I---I	150
Cl---Cl	240
I---Cl	210

(a) - 870 kJ (b) - 390 kJ (c) + 180 kJ (d) + 450 kJ (e) + 1,260 kJ

FREE RESPONSE REVIEW: Thermochemistry

Problem 120:

Reaction	Equation	ΔH°_{298}	ΔS°_{298}	ΔG°_{298}
X	$C(s) + H_2O(g) \rightleftarrows CO(g) + H_2(g)$	+131 kJ mol^{-1}	+134 J mol^{-1} K^{-1}	+91 kJ mol^{-1}
Y	$CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g)$	+41 kJ mol^{-1}	+42 J mol^{-1} K^{-1}	+29 kJ mol^{-1}
Z	$2\,CO(g) \rightleftarrows C(s) + CO_2(g)$	?	?	?

Answer the following questions using the information related to reactions X, Y, and Z in the table above.

(a) For reaction X, write the expression for the equilibrium constant, *Kp* .

(b) For reaction X, will the equilibrium constant, *Kp* , increase, decrease, or remain the same if the temperature rises above 298 K ? Justify your answer.

(c) For reaction Y at 298 K, is the value of *Kp* greater than 1, less than 1, or equal to 1? Justify your answer.

(d) For reaction Y at 298 K, which is larger: the total bond energy of the reactants or the total bond energy of the products? Explain.

(e) Is the following statement true or false? Justify your answer. "On the basis of the data in the table, it can be predicted that reaction Y will occur more rapidly than reaction X will occur."

(f) Consider reaction Z at 298 K.

(i) Is $\Delta S°$ for the reaction positive, negative, or zero? Justify your answer.

(ii) Determine the value of $\Delta H°$ for the reaction.

(iii) A sealed glass reaction vessel contains only CO(*g*) and a small amount of C(*s*). If a reaction occurs and the temperature is held constant at 298 K, will the pressure in the reaction vessel increase, decrease, or remain the same over time? Explain.

Problem 121:

$$N_2(g) + 3\,F_2(g) \rightarrow 2\,NF_3(g) \qquad \Delta H^\circ_{298} = -264 \text{ kJ mol}^{-1};\ \Delta S^\circ_{298} = -278 \text{ J K}^{-1} \text{ mol}^{-1}$$

The following questions relate to the synthesis reaction represented by the chemical equation in the box above.

(a) Calculate the value of the standard free energy change, ΔG°_{298}, for the reaction.

(b) Determine the temperature at which the equilibrium constant, Keq, for the reaction is equal to 1.00.
(assume that ΔH° and ΔS° are independent of temperature.)

(c) Calculate the standard enthalpy change, ΔH°, that occurs when a 0.256 mol sample of $NF_3(g)$ is formed from $N_2(g)$ and $F_2(g)$ at 1.0 atm and 298 K.

The enthalpy change in a chemical reaction is the difference between energy absorbed in breaking bonds in the reactants and energy released by bond formation in the products.

(d) How many bonds are formed when two molecules of NF_3 are produced according to the equation in the box above?

(e) Use both the information in the box above and the table of average bond enthalpies below to calculate the average enthalpy of the F-F bond.

Bond	Average Bond Enthalpy (kJ mol^{-1})
N≡N	946
N–F	272
F–F	?

Problem 122: $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

The combustion of carbon monoxide is represented by the equation above.

(a) Determine the value of the standard enthalpy change, ΔH^o_{rxn}, for the combustion of CO(g) at 298 K using the following information.

$$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g) \qquad \Delta H^\circ_{298} = -110.5 \text{ kJ mol}^{-1}$$

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^\circ_{298} = -393.5 \text{ kJ mol}^{-1}$$

(b) Determine the value of the standard entropy change, ΔS^o_{rxn}, for the combustion of CO(g) at 298 K using the information in the following table.

Substance	S°_{298} ($\text{J mol}^{-1}\text{ K}^{-1}$)
$CO(g)$	197.7
$CO_2(g)$	213.7
$O_2(g)$	205.1

(c) Determine the standard free energy change, ΔG^o_{rxn}, for the reaction at 298 K. Include units with your answer.

(d) Is the reaction spontaneous under standard conditions at 298 K? Justify your answer.

(e) Calculate the value of the equilibrium constant, Keq, for the reaction at 298 K.

AP CHEMISTRY REVIEW: BIG IDEA # 6
EQUILIBRIUM (K, Kp, Ksp, Q, Ka, Kb)

Distinguish chemical formulas as strong, weak, and nonelectrolytes

Identify the formulas of strong and weak acids
Write chemical equations for the ionization of strong and weak acids and bases
Write balanced net ionic equations for precipitation reactions, neutralization reactions and reactions between acids and bases
Describe and interpret the interactive forces between solute and solvent ions and molecules
Construct a visual representation of solute and solvent particles in solution
Apply enthalpy and entropy principles to justify the factors that affect solubility of solutes
Explain solubility data of ionic substances in water and justify with considerations of energy and entropy
Use chemical formulas to predict the solubilities of compounds in water and other solvents.
Calculate molarity and construct visual models to demonstrate different molar concentrations
Write equilibrium and nonequilibrium expressions for K and Q
Explain observations in terms of reversibility of a chemical or physical system under a specified set of conditions
Predict how Q and K change when conditions change
Calculate equilibrium constants using numerical or graphical data
Use stoichiometry and the law of mass action to calculate equilibrium concentrations and partial pressures
Predict the relative concentrations of reactants and products from magnitudes of equilibrium constants
Use le Chatelier's principle and kinetics to predict the relative rates of forward and reverse reactions
Use le Chatelier's principle to predict the direction a reaction will proceed as a result of a given change
Use Le Chatelier's principle to explain the effect a change will have on Q or K
Use Le Chatelier's principle to design a set of conditions that will optimize a desired result
Estimate and calculate quantities such as pH, molar concentrations and various species, and percentage ionization of solutions of strong and weak acids and bases
Write chemical equations that illustrate the Bronsted-Lowry definition of acid and bases
Identify acid- base conjugate pairs
Interconvert Ka and Kb expressions for conjugate pairs
Construct particle representations for the reactions of acids with bases
Write and explain equations for acid-base hydrolysis of salt solutions
Justify with examples why chemical structure affects acid- base behavior
Given a mixture of strong and weak acids and bases, predict the reaction that will occur, write a net ionic equation and identify what chemical species will be present at equilibrium
Calculate the pH of a buffer, given the concentrations of its components
Calculate the concentrations needed to obtain a desired pH and buffer capacity
Tell whether the conjugate acid or the conjugate base will predominate in solution at a given pH, given the pKa of an acid
Identify a buffer from its components or its behavior in solution
Explain how a buffer works to resist a change in pH
Interpret weak acid-strong base titration data to determine the equivalence point, the pH at the equivalence point, the concentration of the weak acid, and the pKa of the weak acid
Predict and rank the solubilities of various salts, given their solubility product constants
Explain and justify the factors that influence the solubility of salts

Equilibrium expression/law of mass action/equilibrium constant expression

$$aA + bB \rightleftharpoons cC + dD$$

(or Kc) $K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ ** do NOT include pure solids or pure liquid (or units!)!

Kp- do NOT use brackets, use P or ( ), include gas ONLY!
If mol gas product = mol gas reactant then Kp = Kc

When system is NOT at equilibrium but you are given INITIAL values and need to determine which way system is shifting:

Use REACTION QUOTIENT:

$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ ** do NOT include pure solids or pure liquid (or units!)!

Comparing Q to K:
K > Q reaction is driven right, products favored
K < Q reaction is driven left, reactants favored
K = Q reaction is at equilibrium

Le Chatlier: If a system is stressed, the system will shift in order to reduce the stress.

Types of stress: Change in concentration, change in temperature, change in pressure or volume (gas ONLY)

$$A(s) + 2B(g) \rightleftharpoons 3C(g) + D(aq) + \text{heat}$$

Stress: Add heat

Remove heat

Add pressure

Remove pressure

Add B

Remove B

Add C

Remove C

Solubility Equilibrium: Used for salts that is generally considered insoluble.

Ksp (s) $\rightleftharpoons$ + ion (aq) + - ion (aq)

K < Q precipitate forms!

K > Q more dissolving occurs!

"molar solubility" is code for "x" There is NEVER a coefficient in Ksp equilibrium expressions since Ksp reactant is always a solid.
Common ion effect and uncommon ion effect

MULTIPLE CHOICE REVIEW: Equilibrium

_____123. $H_2(g) + Br_2(g) \rightleftharpoons 2\ HBr(g)$

At a certain temperature, the value of the equilibrium constant, K, for the reaction represented above is 2.0×10^{5}. What is the value of K for the reverse reaction at the same temperature?

(a) -2.0×10^{-5} (b) 5.0×10^{-6} (c) 2.0×10^{-5} (d) 5.0×10^{-5} (e) 5.0×10^{-4}

_____124. In a saturated solution of $Zn(OH)_2$ at 25ºC, the value of $[OH^-]$ is 2.0×10^{-6} M. What is the value of the solubility-product constant, Ksp, for $Zn(OH)_2$ at 25ºC?

(a) 4.0×10^{-18} (b) 8.0×10^{-18} (c) 1.6×10^{-17} (d) 4.0×10^{-12} (e) 2.0×10^{-6}

_____125. $2\ K + 2\ H_2O \rightarrow 2\ K^+ + 2\ OH^- + H_2$

When 0.400 mole of potassium reacts with excess water at standard temperature and pressure as shown in the equation above, the volume of hydrogen gas produced is

(a) 1.12 liters (b) 2.24 liters (c) 3.36 liters (d) 4.48 liters (e) 6.72 liters

_____126. The solubility of CuI is 2×10^{-6} molar. What is the solubility product constant, K_{sp}, for CuI?

(a) 1.4×10^{-3} (b) 2×10^{-6} (c) 4×10^{-12} (d) 2×10^{-12} (e) 8×10^{-18}

_____127. $MnS(s) + 2\ H^+ \rightleftharpoons Mn^{2+} + H_2S(g)$

At 25 °C the solubility product constant, Ksp, for MnS in 5×10^{-15} and the acid dissociation constants K_1 and K_2 for H_2S are 1×10^{-7} and 1×10^{-13}, respectively. What is the equilibrium constant for the reaction represented by the equation above at 25 °C?

(a) $\frac{1 \times 10^{-13}}{5 \times 10^{-15}}$ (b) $\frac{5 \times 10^{-15}}{1 \times 10^{-7}}$ (c) $\frac{1 \times 10^{-7}}{5 \times 10^{-20}}$ (d) $\frac{5 \times 10^{-15}}{1 \times 10^{-20}}$ (e) $\frac{1 \times 10^{-20}}{5 \times 10^{-15}}$

_____128. How many moles of NaF must be dissolved in 1.00 liter of a saturated solution of PbF_2 at 25 °C to reduce the $[Pb^{2+}]$ to 1×10^{-6} molar? (K_{sp} of PbF_2 at 25 °C = 4.0×10^{-8})

(a) 0.020 mole (b) 0.040 mole (c) 0.10 mole (d) 0.20 mole (e) 0.40 mole

_____129. What is the molar solubility in water of Ag_2CrO_4? (The Ksp for Ag_2CrO_4 is 8×10^{-12}.)

(a) 8×10^{-12} M (b) 2×10^{-12} M (c) (4×10^{-12} M)1/2 (d) (4×10^{-12} M)1/3 (e) (2×10^{-12} M)1/3

FREE RESPONSE REVIEW: Equilibrium

Problem 130:

Several reactions are carried out using AgBr, a cream-colored silver salt for which the value of the solubility product constant, *Ksp* , is 5.0×10^{-13} at 298 K.

(a) Write the expression for the solubility-product constant, *Ksp* , of AgBr.

(a) Calculate the value of $[Ag^+]$ in 50.0 mL of a saturated solution of AgBr at 298 K.

(b) A 50.0 mL sample of distilled water is added to the solution described in part (b), which is in a beaker with some solid AgBr at the bottom. The solution is stirred and equilibrium is reestablished. Some solid AgBr remains in the beaker. Is the value of [Ag+] greater than, less than, or equal to the value you calculated in part (b) ? Justify your answer.

(c) Calculate the minimum volume of distilled water, in liters, necessary to completely dissolve a 5.0 g sample of AgBr(*s*) at 298 K. (The molar mass of AgBr is 188 g mol^{-1}.)

(d) A student mixes 10.0 mL of 1.5×10^{-4} *M* AgNO3 with 2.0 mL of 5.0×10^{-4} *M* NaBr and stirs the resulting mixture. What will the student observe? Justify your answer with calculations.

(e) The color of another salt of silver, AgI(*s*), is yellow. A student adds a solution of NaI to a test tube containing a small amount of solid, cream-colored AgBr. After stirring the contents of the test tube, the student observes that the solid in the test tube changes color from cream to yellow.

(i) Write the chemical equation for the reaction that occurred in the test tube.

(ii) Which salt has the greater value of *Ksp* : AgBr or AgI ? Justify your answer.

Problem 131:

The compound butane, C_4H_{10} , occurs in two isomeric forms, *n*-butane and isobutane (2-methyl propane). Both compounds exist as gases at 25°C and 1.0 atm.

(a) Draw the structural formula of each of the isomers (include all atoms). Clearly label each structure.

(b) On the basis of molecular structure, identify the isomer that has the higher boiling point. Justify your answer.

The two isomers exist in equilibrium as represented by the equation below.

$$n\text{-butane}(g) \rightleftharpoons \text{isobutane}(g) \qquad K_c = 2.5 \text{ at } 25°C$$

Suppose that a 0.010 mol sample of pure *n*-butane is placed in an evacuated 1.0 L rigid container at 25°C.

(a) Write the expression for the equilibrium constant, *Kc* , for the reaction.

(b) Calculate the initial pressure in the container when the *n*-butane is first introduced (before the reaction starts).

(c) The *n*-butane reacts until equilibrium has been established at 25°C.

(i) Calculate the total pressure in the container at equilibrium. Justify your answer.

(ii) Calculate the molar concentration of each species at equilibrium.

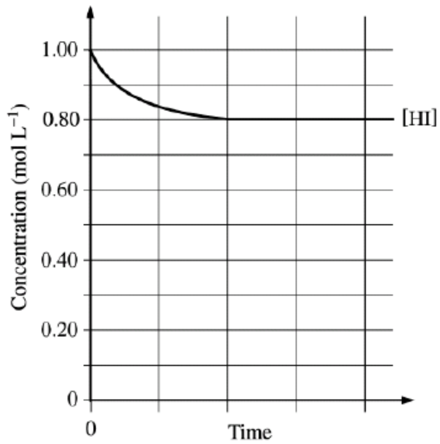
Suppose that in another experiment a 0.010 mol sample of pure isobutane is placed in an evacuated 1.0 L rigid container and allowed to come to equilibrium at 25°C.

(d) Calculate the molar concentration of each species after equilibrium has been established.

Problem 132:

After a 1.0 mole sample of HI(g) is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of HI(g) as a function of time is shown below.

$$2\,HI(g) \leftrightarrows H_2(g) + I_2(g)$$



(a) Write the expression for the equilibrium constant, Kc , for the reaction.

(b) What is [HI] at equilibrium?

(c) Determine the equilibrium concentrations of $H_2(g)$ and $I_2(g)$.

(d) On the graph above, make a sketch that shows how the concentration of $H_2(g)$ changes as a function of time.

(e) Calculate the value of Kc, the equilibrium constant, at 700. K.

(f) At 1,000 K, the value of Kc for the reaction is 2.6×10^{-2}. In an experiment, 0.75 mole of HI(g), 0.10 mole of $H_2(g)$, and 0.50 mole of $I_2(g)$ are placed in a 1.0 L container and allowed to reach equilibrium at 1,000 K. Determine whether the equilibrium concentration of HI(g) will be greater than, equal to, or less than the initial concentration of HI(g). Justify your answer.

AP CHEMISTRY REVIEW: Acids and Bases

Acid and Base Definitions:

Arrhenius acid: a substance that, when dissolved in water, increases the concentration of H^+ ions in solution
Arrhenius base: a substance that, when dissolved in water, increases the concentration of OH^- ions in solution.

**Memorize Strong Acids: "the big 6" $HClO_4$ (perchloric acid), HNO_3 (nitric acid), H_2SO_4 (sulphuric acid)
HCl, HBr, HI (hydrochloric, hydrobromic, and hydroiodic acids),
Strong Bases: Group 1 hydroxides (LiOH, NaOH, KOH…)
Group 2 (ledge down) $Ba(OH)_2$, $Sr(OH)_2$

Bronsted-Lowry Definitions:

Bronsted-Lowry acid: a substance that transfers a proton to another substance (H^+ donor)
Bronsted-Lowry base: a substance that accepts a proton from another substance (H^+ acceptor)

Bronsted-Lowry Conjugate Pairs:
differ only by the presence (or absence) of a proton, which classifies them as conjugate acid-base pairs.
Strong acid has a weak conjugate base (and vice versa)
Weak acid has a strong conjugate base (and vice versa)

Amphoterism: a substance that acts as an acid in some reactions and a base in others.

Neutralization Reaction: Acid + Base → Salt + Water
Hydrolysis Reaction: Salt + water → Acid + Base

$Kw = [H^+][OH^-] = [H3O^+][OH^-] = 1.0 \times 10^{-14}$
Neutral $pH = 7$, $[H^+] = [OH^-] = 1.0 \times 10^{-7}$
Acidic $pH < 7$, $[H^+] > [OH^-]$
Basic $pH > 7$ $[H^+] < [OH^-]$

<u>pH scale:</u> "power of hydrogen ion" in solution
$pH = -\log[H^+]$
$[H^+] = 10^{-pH}$

pOH scale: "power of hydroxide ion"
$pOH = -\log[OH^-]$
$[OH^-] = 10^{-pOH}$

<u>Weak Acids and Weak Bases:</u>

Ka - acid ionization constant
$HA \rightleftharpoons H^+ + A^-$

$$Ka = \frac{[H^+][A^-]}{[HA]}$$

Kb – base ionization constant
$WB \rightleftharpoons HB^+ + OH^-$

$$Kb = \frac{[OH^-][HB^+]}{[WB]}$$

** if problem requires quadratic equation, Neglect x!

Hydrolysis: use when @ equivalence point with WA/SB and WB/SA problems!
SA + WB ex. $MgCl_2$
H^+

WA + SB ex. NaF
OH^-

$K_w = K_aK_b$

Acid – Base Titration:

A **titration** is a **volumetric analysis** because you carefully measure the volume of **titrant**, dispensing it from a **buret**. When you have added just enough titrant to completely react with the sample, you have reached the equivalence point. At the equivalence point you have equal mole of the substance you started with and the substance you were adding.

If the proper Indicator was chosen the equivalence point is usually apparent because of the color change .

SA + SB **WA + SB** **WB + SA**

At start:

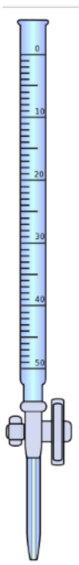
After 1st drop:

At half Equivalence:

Just past half Equivalence:

At Equivalence:

Past Equivalence:



In AP Chemistry, it is safe to assume that p(anything) = -log(anything)

$pK_a = -\log K_a$ $pK_b = -\log K_b$

MULTIPLE CHOICE REVIEW: Acids and Bases

_____133. If the acid dissociation constant, K_a, for an acid HA is 8 x 10^{-4} at 25 °C, what percent of the acid is dissociated in a 0.50-molar solution of HA at 25 °C?

(a) 0.08% (b) 0.2% (c) 1% (d) 2% (e) 4%

_____134. $5\ Fe^{2+} + MnO_4^{-} + 8\ H^{+} \rightleftharpoons 5\ Fe^{3+} + Mn^{2+} + 4\ H_2O$

In a titration experiment based on the equation above, 100.0 milliliters of an acidified Fe^{2+} solution requires 14.0 milliliters of standard 0.050-molar MnO_4^{-} solution to reach the equivalence point. The concentration of Fe^{2+} in the original solution is

(a) 0.0035 M (b) 0.0070 M (c) 0.035 M (d) 0.070 M (e) 0.14 M

_____ 135. $H_2C_2O_4 + 2\ H_2O \rightleftharpoons 2\ H_3O^{+} + C_2O_4^{2-}$

Oxalic acid, $H_2C_2O_4$, is a diprotic acid with K_1 = 5 x 10^{-2} and K_2 = 5 x 10^{-5}. Which of the following is equal to the equilibrium constant for the reaction represented above?

(a) 5 x 10^{-2} (b) 5 x 10^{-5} (c) 2.5 x 10^{-6} (d) 5 x 10^{-7} (e) 2.5 x 10^{-8}

_____ 136. What is the H^{+}(aq) concentration in 0.05 M HCN (aq) ? (The K_a for HCN is 5.0 x 10^{-10})

(a) 2.5 x 10^{-11} (b) 2.5 x 10^{-10} (c) 5.0 x 10^{-10} (d) 5.0 x 10^{-6} (e) 5.0 x 10^{-4}

_____137. At 25^{O}C, aqueous solutions with a pH of 8 have hydroxide ion concentration, [OH^{-}], of

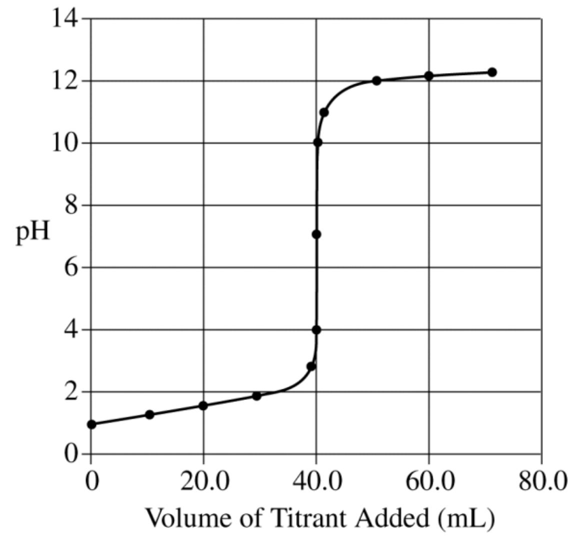
(a) 1 x 10^{-14} M (b) 1 x 10^{-8} M (c) 1 x 10^{-6} M (d) 1 M (e) 8 M

_____ 138. A 0.20-molar solution of a weak monoprotic acid, HA, has a pH of 3.00. The ionization constant of this acid is

(a) 5.0 x 10^{-7} (b) 2.0 x 10^{-7} (c) 5.0 x 10^{-6} (d) 5.0 x 10^{-3} (e) 2.0 x 10^{-3}

FREE RESPONSE REVIEW: Acids and Bases

Problem 139:



A solution of 0.100 *M* HCl and a solution of 0.100 *M* NaOH are prepared. A 40.0 mL sample of one of the solutions is added to a beaker and then titrated with the other solution. A pH electrode is used to obtain the data that are plotted in the titration curve shown above.

(a) Identify the solution that was initially added to the beaker. Explain your reasoning.

(b) On the titration curve above, circle the point that corresponds to the equivalence point.

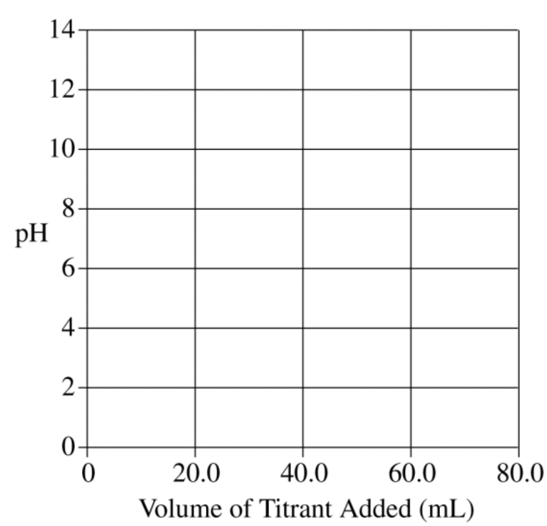
(c) At the equivalence point, how many moles of titrant have been added?

(d) The same titration is to be performed again, this time using an indicator. Use the information in the table below to select the best indicator for the titration. Explain your choice.

Indicator	pH Range of Color Change
Methyl violet	0 – 1.6
Methyl red	4 – 6
Alizarin yellow	10 – 12

(e) What is the difference between the equivalence point of a titration and the end point of a titration?

(f) On the grid provided below, sketch the titration curve that would result if the solutions in the beaker and buret were reversed (i.e., if 40.0 mL of the solution used in the buret in the previous titration were titrated with the solution that was in the beaker).



Problem 140:

A buffer solution contains 0.40 mole of formic acid, HCOOH, and 0.60 mole of sodium formate, HCOONa, in 1.00 liter of solution. The ionization constant, K_a, of formic acid is 1.8×10^{-4}.

(a) Calculate the pH of this solution.

(b) If 100. millilitres of this buffer solution is diluted to a volume of 1.00 liter with pure water, the pH does not change. Discuss why the pH remains constant on dilution.

(c) A 5.00 milliliter sample of 1.00 molar HCl is added to 100. milliliters of the original buffer solution. Calculate the $[H_3O^+]$ of the resulting solution.

(d) A 800.–milliliter sample of 2.00–molar formic acid is mixed with 200. milliliters of 4.80–molar NaOH. Calculate the $[H_3O^+]$ of the resulting solution.

AP Chemistry Equations & Constants

Throughout the test the following symbols have the definitions specified unless otherwise noted.

L, mL	= liter(s), milliliter(s)	mm Hg	= millimeters of mercury
g	= gram(s)	J, kJ	= joule(s), kilojoule(s)
nm	= nanometer(s)	V	= volt(s)
atm	= atmosphere(s)	mol	= mole(s)

ATOMIC STRUCTURE

$E = h\nu$

$c = \lambda\nu$

E = energy

ν = frequency

λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}

Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}

Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$K_c = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$, where $a\,A + b\,B \rightleftarrows c\,C + d\,D$

$K_p = \dfrac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$

$K_a = \dfrac{[H^+][A^-]}{[HA]}$

$K_b = \dfrac{[OH^-][HB^+]}{[B]}$

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

$= K_a \times K_b$

$\text{pH} = -\log[H^+]$, $\text{pOH} = -\log[OH^-]$

$14 = \text{pH} + \text{pOH}$

$\text{pH} = \text{p}K_a + \log\dfrac{[A^-]}{[HA]}$

$\text{p}K_a = -\log K_a$, $\text{p}K_b = -\log K_b$

<u>Equilibrium Constants</u>

K_c (molar concentrations)

K_p (gas pressures)

K_a (weak acid)

K_b (weak base)

K_w (water)

KINETICS

$\ln[A]_t - \ln[A]_0 = -kt$

$\dfrac{1}{[A]_t} - \dfrac{1}{[A]_0} = kt$

$t_{1/2} = \dfrac{0.693}{k}$

k = rate constant

t = time

$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{total} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, $R = 8.314\ \text{J mol}^{-1}\ \text{K}^{-1}$
$= 0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1}$
$= 62.36\ \text{L torr mol}^{-1}\ \text{K}^{-1}$
1 atm = 760 mm Hg
= 760 torr
STP = $0.00\,^\circ$C and 1.000 atm

THERMOCHEMISTRY/ ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= -RT \ln K$$

$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1\ \text{volt} = \frac{1\ \text{joule}}{1\ \text{coulomb}}$$

PERIODIC TABLE OF THE ELEMENTS

1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

*Lanthanide Series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

1. B
2. C
3. A
4. C
5. C
6. D
7. D
8. A
9. a.) CH_2
   b.) 56 g/mol
   c.) C_4H_8
10. C
11. 91.42
12. D
13. C
14. B
15. D
16. C
17. D
18. B
19. E
20. B
21. D
22. a.) $8s^1$
   b.) metal due to one valence electron
   c.) largest due to 8 shells compared to other members of group who have fewer shells d.) Q^{+1}
   e.) (i) Q^{+1} CO_3^{-2} Q_2CO_3 (ii) yes because group 1 ions are soluble, always soluble in H_20
23. a.) Mg is smaller therefore stronger hold on its valence electrons
   b.) 1st and 2nd are valence electrons, 3rd is a core electron therefore closer to nucleus therefore more difficult to remove
   c.) Na is a metal with mobile valence electrons therefore capable of carry electricity NaCl is an ionic compound with no moving particles
   d.) the 1st electron lost for Al is in p orbitals which penetrates nucleus less than s orbital of Mg valence electron
   e.) radius increases because all have same number of shells but Al has most protons therefore strongest attraction to valence electrons therefore smallest radius
24. C
25. B
26. C
27. C
28. D
29. C

30. B
31. E
32. A
33. C
34. D
35. A) 2.59 atm
    B) 26.69 mol total
36. A) 741 torr
    B) 0.0073 mol
    C) 0.4964 g H2O2
    D) 7.14%
37. A) 14.01 atm Ptotal
    B) 0.714
    C) 5.01 g/L
    D) H2: 1.50 mol/3.0 mol= 0.5, H2O: 1.0 mol/ 3.0 mol= 0.33, Ar: 0.5 mol/ 3.0 mol= 0.167
38. A
39. C
40. D
41. B
42. C
43. D
44. B
45. A
46. D
47. B
48. B
49. C
50. C
51. C
52. C
53. A) Solid K metal ∴ mobile ve-. KNO3 ionic solid ∴ no moving charged particles.

    B) SbCl3: 5+21= 26/2= 13, Trigonal Planar. SbCl5: 5+35= 40/2= 20, Trigonal Bipyramidal ∴ symmetrical.

    C) CBr4 has more e- ∴ more polarizable ∴ stronger IMFs ∴ requires more E to break CBr4 bonds ∴ boils at a higher temp.

    D) I2 has a greater boiling point due to more e- ∴ more polarizable ∴ stronger IMFs. Br2 is a smaller atom ∴ Fc between ve- is stronger ∴ greater BE.
54. A) Both are nonpolar, but F2 has 18 ve- and I2 has 106 ve- ∴ more polarizable ∴ stronger IMFs.

B) Both are ionic ∴ $Fc \propto Q^{+}Q^{-}/r^{2}$. NaF: +1-1 ∴ small, CsCl +1-1 ∴ large.

C) I has extra e- pairs on its central atom.

D) N-H is large difference in eneg ∴ creates extreme dipoles that have the ability to hydrogen bond.

55. B> A=F > C=D=E
56. D
57. B
58. C
59. $Na_2S_2O_{3(aq)}$ because we need 0.54V to be the highest # so it gets reduced ∴ $S_2O_3^{-2}$ needs to be our reactant.
60. + 0.46 V
61. 2 $S_2O_{3-}^{2}$ + I_2 → 2I- + $S_4O_5^{-2}$
62. omit
63. a) H_2O b) 9.459 g H_2S c) 1.11 g Al_2S_3
64. A
65. A
66. A
67. B
68. D
69. B
70. B
71. C
72. B
73. B
74. C
75. D
76. A
77. B
78. A
79. D
80. A
81. B
82. D
83. D
84. E
85. D
86. B
87. C
88. D
89. C
90. D

91. B
92. B

93a. cathode: 2e- + Ca^{+2} → Ca

anode:$2Cl^-$ → Cl_2 + $2e^-$

93b. .506g Ca and .885g Cl_2
93c. Reactions would be the same because group 2 metals get Reduced in presence of H^2O and Cl^- would be oxidized instead of O^{-2}
94. B
95. B
96. D
97. A
98. A
99. E
100a. $2Ag^+$ + Cd → 2Ag + Cd^{+2} E= 1.20V
100b. Anions flow to cell containing Cd anode, this is in order to maintain neutrality as Cd^{+2} ions are dropped into solution, No^{-3} ions will neutralize them.
100c. Ag+ is a reactant therefore according to Le Chatelier's principle there will be an increase in collisions and reaction will drive right therefore voltage will increase.
100d. AgCl will precipitate therefore Ag^+ concentration will decrease therefore the reaction will shift in reverse therefore the voltage will decrease.
100e.Adding H_2O has 2x the impact on Ag^+ as it does on Cd^{+2} therefore the Ag^+ decreases significantly therefore driving the reaction left therefore there is a decrease in voltage
101a.Keq = $4.2E^{17}$ because K>1 there will be more products created than reactants therefore spontaneous therefore Ered is positive.
101b. As T increases so does E
101c. Reaction would drive right based on Le Chatelier's equation if you decrease the product the reaction shifts to replace the lost ions therefore to the right which increases E.
101d.E = 0 change in G=0 K=1 Dead Battery
102a. 2Cl- → Cl_2 +$2e^-$
102b. FeCl2 Fe^{+2}
102c. Fe^{+2} + $2Cl^-$ → Cl_2 + Fe
102d. 0.231 L
102e. 2.30 c/s
103. D
104. B
105. Top row left to right: D, C, F Bottom row left to right: E, B, A
106.
- 1a. Zero with respect to A
- 1b. 2nd order w/ respect to B
- 2. rate= $K[B]^2$
- 3. K = .577 mol^{-1} s^{-1}

4. 0.0313 m/s
5. .00214 m/s
6a. No because the stoichiometry of the reaction is not correct.
6b. Intermediate because it is formed in one step and used up in the next.

107. a) 1.0 atm
b) 2nd order because reciprocal plot is straight line
c) 2.5 L $mol^{-1}s^{-1}$

108. Omit
109. omit
110. Same b/c @ same temp
111. H_2 because it's the lightest
112. System is getting colder, heat is being released, exothermic
113. $1.3x10^3$
114. 9.47 KJ
115. H-Bonding
116. H_2
117. A
118. C
119. B
120. a). Kp= (Pco) (P_{H2})/(P_{H2O})
b). According to Le Chatlier if temp increases the reaction shifts right in the endothermic direction so the number of products increases so the Kp increases.
c). Less than 1
d). The bond energy of the products is greater
e). False Thermodynamics does not affect how fast a reaction occurs.
f). i). Change in Entropy is negative 2 mol of gas to 1 mol of gas and 1 mol of solid
ii). -172 KJ/Mol
iii). Fewer moles present, Pressure decreases.

121. a). -181 KJ/mol
b). 950. K
c). -33.8 KJ
d). 6 bonds
e). 141 KJ/mol

122. a) - 504 kJ/mol
b) -86.55 J/mol K
c) - 478 kJ/mol
d) Yes, G is negative therefore spontaneous
e) 1.21

123. B
124. A
125. D
126. C
127. D
128. D

129. E
130. a). Ksp= [Ag+] [Br-]
b). $7.1x10^{-7}$
c). [Ag+] is the same more Ag dissolves until the solution is saturated again
d). $1E^{-6}$ mole will ppt.
e). i). NaI+AgBr ---> AgI+NaBr
ii). AgBr has a greater Ksp since AgI ppt.
131. a).
b). n-butane its shape has stronger IMF's
a). Kc= [Isobutane]/[n-butane]
b). P=0.245 atm
c). i). .245atm
ii). [isobutane]= 0.00714 [n-butane]= 0.00286
d). Same answers ^
132. a) $K = [H_2] [I_2] / [HI]^2$
b) 0.80 M
c) 0.01 M
d) line on graph at .10 M
e) 0.0156
f) Greater because K<Q therefore reaction will shift left making more reactant
133. E
134. C
135. C
136. D
137. C
138. C
139.) a.) .100 M HCl because initial pH was1 therefore strong acid present
b.) equivalence point is approx. 7 at 40 ml
c.) mv= mol (.100M)(.040L) = 0.004mol
d.) methyl red because you want an indicator that changes color at equivalence point
e.) mol acid = mol base when indicator changes color **f.) *graph* goes from 12 to 1 equivalence point at 7
140.) a.) pH = 3.92
b.) pH is determined by ratio WA:CB therefore ratio is constant
c.) [H+] = $1.23E^{-4}$ M
d.) [H+] = 1.20 E^{-4} M