**Unit 1 Measurements**

Things to know: scientific method, qualitative/quantitative data, SI measurements (m, L, g), ways to find volume, converting with prefixes, *density*, accuracy vs. precision, *% error*, significant figures, scientific notation, atom/element/compound, chemical/physical properties, chemical/physical changes, classification of matter (mixture/pure substance and homogeneous/heterogeneous), periodic table (periods/groups/families, metals/nonmetals/metalloids)

\**items in italics is for honors only*

Objective

2.2.2—Analyze the evidence of chemical change.

Problems

1. List the four indicators of a chemical change:
2. Define precipitate. How can you determine if something is a precipitate or not?
3. What would happen if a burning splint was exposed to:
   1. Oxygen
   2. Hydrogen
   3. Carbon dioxide
4. What would happen if carbon dioxide was bubbled through lime water?
5. What piece of safety equipment must be used in all laboratory experiments?
6. When using a Bunsen burner what precautions should be made?
7. Exothermic reaction \_\_\_\_\_\_\_\_\_\_\_\_ energy, ∆H is \_\_\_\_ and energy is shown on the \_\_\_\_\_\_ side.
8. Endothermic reaction \_\_\_\_\_\_\_\_\_\_\_ energy, ∆H is \_\_\_\_and energy is shown on the \_\_\_\_\_\_ side.
9. Label each of the following as a physical or chemical property and give a brief reason
   1. Flammability c. Can neutralize a base
   2. Density d. Boiling Point
10. Label each of the following as a physical or chemical change
    1. A tire is inflated with air. c. Water is added to red soln, it turns pink.
    2. Food is digested in the stomach. d. Water is heated and changed into steam.

**Unit 2 Atoms**

Things to know: Dalton’s atomic theory fallacies, Thompson’s plum pudding model, Rutherford’s gold foil experiment, planetary model, atom’s particles names and locations and masses, isotopes and the different notations, nucleus, atomic number, mass number, average atomic mass, mole, Avogadro’s number, molar mass, conversions (using mole, Avogadro’s number and molar mass), nuclear chemistry why we have it, types of particles (alpha, beta, gamma), balancing nuclear reactions, nuclear fission/fusion

Objectives

1.1.1—Analyze the structure of atoms, isotopes, and ions.

1.1.4-Explain the process of radioactive decay using nuclear equations and half-life.

2.2.4—Analyze the stoichiometric relationships inherent in a chemical reaction.

Problems

1. List the three particles and their charges, locations, and relative masses.
2. Isotopes differ in the number of \_\_\_\_\_\_\_\_\_\_\_ not \_\_\_\_\_\_\_\_\_\_\_ (which always must stay the same as they indicate the atomic number!)

Ex. AZX Ex. F-20

Z means \_\_\_\_\_\_\_\_\_\_\_\_ Atomic Number \_\_\_\_\_\_\_\_\_\_\_\_

A means \_\_\_\_\_\_\_\_\_\_\_\_ Mass Number \_\_\_\_\_\_\_\_\_\_\_\_

1. Fill out the chart below for each of the three isotopes

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Isotope | Electrons | Protons | Mass Number | Atomic Number | Neutrons |
| 238U |  |  |  |  |  |
| 16O2- |  |  |  |  |  |
| 23Na+1 |  |  |  |  |  |

1. What is the difference between average atomic mass, isotopic mass, and mass number?
2. Balancing Nuclear Reactions (remember Law of Conservation of Mass) and label fission or fusion
   1. \_\_\_\_\_\_\_\_\_🡪3717Cl + 11H ­­­­­­­­­­­­­ b. 6430Zn + \_\_\_\_\_\_\_\_\_ 🡪6429Cu + 10n
3. Give the symbol for the following
   1. Alpha b. Beta c. Gamma
4. Which type of decay listed above is the strongest? The weakest? Explain why.
5. Decay is a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ event, independent of other energy influences.
6. Balance the following nuclear reactions and label alpha decay or beta decay
   1. 220Rn 🡪4He + \_\_\_\_\_ b. 216Po 🡪0-1e + \_\_\_\_\_
7. Tritium, 3H, has a half-life of 12.3 years. How long would it take for a 40.0g sample to decay down to 1.25g?
8. Fe-61 has a half-life of 6.00min. Of a 100.0g sample, how much will remain after 18.0min?
9. After 20.0 days, a 120kg sample of Bi-210 decays down to just 7.5kg. What is its half-life?
10. What is the molar mass of CO2?
11. What is the molar mass of (NH4)2S • 3H2O?
12. How many grams are in 3.0 moles of H2SO4?
13. How many molecules are in 64 grams of O2?
14. How many moles are in 84.2 grams of CO2?
15. How many moles are in 3.04x1023 molecules of H2?
16. How many grams are in 4.59x1025 particles of NaCl?

**Unit 3 Electron Arrangement and Periodicity**

Things to know: Bohr’s model, line emission spectrum and why it happens, wavelength vs. frequency vs. energy, c *= λν*, *E = hν*, Heisenberg’s uncertainty principle, Schrödinger’s wave equation, de Broglie, quantum numbers, sublevels, orbitals for each sublevel, Aufbau principle, Hund’s rule, Pauli exclusion principle, orbital notation, electron configuration, noble gas configuration, valence electrons, Lewis dot diagram, history of periodic table,Mendeleev, Moseley, modern periodic law, family names, three things all families have in common, trends, isoelectronic, cations vs. anions, \**items in italics is for honors only*

Objectives

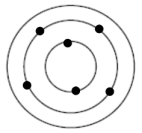
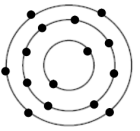
1.1.2—Analyze an atom in terms of the location of electrons.

1.1.3—Explain the emission of electromagnetic radiation in spectral form in terms of Bohr’s model.

1.3.1—Classify the components of a periodic table.

1.3.2.—Infer the physical properties of an element based on its position on the periodic table.

Problems

1. According to Bohr’s model of the atom, where is the only place electrons can be located?
2. Describe the electron cloud.
3. Use the diagram below to draw the 7Li isotope.
4. Label each diagram with the element (***name and symbol***) they represent.
   1. 
   2. 
5. Define quanta.
6. Describe the difference between the ground state and an excited state. How do electrons move from one to the other?
7. Define photon.
8. What is the relationship between
   1. Wavelength and frequency b. energy and frequency
9. What is the wavelength of a photon emitted when the electron falls from the third energy level to the second energy level? What type of electromagnetic radiation is it?
10. What is the wavelength of a photon emitted when the electron falls from the sixth energy level to the third energy level? What type of electromagnetic radiation is it?
11. Describe the wave/particle duality of electrons.
12. Niels Bohr produced a model of the hydrogen atom based on experimental observations. This model indicated that:
    1. An electron \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ the nucleus only in fixed energy ranges called \_\_\_\_\_\_\_\_.
    2. An electron can neither gain nor lose energy \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ this orbital, but could move up or down to another orbit.
    3. The lowest energy orbit is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ to the nucleus.
13. Define groups.
14. Name three things main group elements in the same family have in common.
15. Reactivity \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ down a group for metals and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ for nonmetals.
16. Define periods.
17. Where are the following elements located on the periodic table:
    1. Metals b. Nonmetals c. Metalloids
18. Label the following elements as a metal, nonmetal, or metalloid
    1. Oxygen b. lead c. silicon d. magnesium e. boron f. neon
19. Give the location of the following families on the periodic table
    1. Representative/main group
    2. Alkali metals
    3. Alkaline earth metals
    4. Halogens
    5. Noble gases
    6. Transition elements
20. Define atomic radius.
21. Define ionic radius.
22. What is the group and period trend for atomic radius?
23. Put the following in order by decreasing atomic radius: Al, Na, S, K and explain why.
24. Which of the following has the most metallic character: Al, Na, S, or K?
25. Write the electron configuration for Li and F. Explain which will lose electrons and which will gain electrons to become stable.
26. The more metallic an element is the \_\_\_\_\_\_\_\_\_ the ionization energy, \_\_\_\_\_\_\_\_\_ the electron affinity, and the \_\_\_\_\_\_\_\_\_ the electronegativity.
27. Name the four sublevels and the area it represents on the periodic table.
28. Write the orbital notation for the following:
    1. Sulfur
    2. Nickel
29. Write the electron configuration for the following elements
    1. Boron
    2. Copper
30. Identify the element represented by the configurations below
    1. 1s22s22p4 b. [Ar]4s23d104p3
31. Determine the number of valence electrons in the following configurations
    1. 1s22s22p63s23p64s23d104p5 b. [Kr]5s1
32. List the number of valence electrons for each of the following
    1. Sodium b. Nitrogen c. Bromine
33. Using the configurations below, determine the number of electrons lost or gained and the oxidation number it will form
    1. 1s22s22p63s23p64s23d104p65s2 b. 1s22s22p63s23p4
34. Define ionization energy.
35. What is the general trend for ionization energy? Explain the reasoning.
36. Put the following elements in order of decreasing ionization energy: Rb, Al, S, Mg
37. Define electronegativity.
38. What is the general trend for electronegativity? Explain the reasoning.
39. Put the following elements in order of increasing electronegativity energy: F, B, N, Li

**Unit 4Ionic and Metallic Bonding**

Things to know: bond, octet rule, types of bonds (what electrons do in each), metallic bonding(what electrons do and properties), ionic bonding(what electrons do and properties), writing formulas and naming ionic compounds (molecular ions, binary ionic compounds, polyatomic compounds, and transition/post transition metal compounds), percent composition, hydrates

Objectives

1.2.1—Compare the relative strengths of bonds

1.2.2—Infer the type of bond and chemical formula between atoms

1.2.5—Compare the properties of ionic, covalent, metallic, and network compounds

1.2.4—Interpret the name and formula of compounds using IUPAC

2.2.5—Analyze quantitatively the composition of a substance

Problems

1. Describe metallic bonds.
2. How are ions formed? Which arrangements are stable?
3. Explain difference between cation and anion.
4. Give the ionic charge for the following groups:
   1. Group 1 b. Group 2 c. Group 13 d. Group 15 e. Group 16 f. Group 17
5. What types of elements form ionic compounds? Explain their electronegativity differences.
6. Predict the chemical formulas of compounds using Lewis structures.
   1. Potassium and Sulfur b. Magnesium and Oxygen
7. Give four properties of ionic compounds and explain why they have these properties.
8. Give six properties of metallic compounds and explain why they have these properties.
9. Write the formula for the following:
   1. Magnesium fluoride d. Calcium nitride
   2. Sodium carbonate e. Ammonium phosphate
   3. Copper (III) bromide f. Tin (IV) oxide
10. Write the names for the following formulas:
    1. FeS d. VBr3
    2. NH4NO2 e. Ba(OH)2
    3. Al2S3 f. (NH4)3P
11. Calculate the percent by mass of water in lithium chromate dihydrate.

**Unit 5Covalent Bonding**

Things to know: properties of covalent bonding (polar and nonpolar), diatomic molecules, bond-length vs. bond energy, electronegativity to determine bond type, Lewis structures, polarity of molecule, resonance, geometry, VESPR theory, intermolecular forces (IMF), writing formulas and naming covalent molecules (binary and acids), *oxidation numbers*, empirical formula, molecular formula

Objectives

1.2.1—Compare the relative strengths of bonds

1.2.2—Infer the type of bond and chemical formula between atoms

1.2.3—Compare inter- and intra- particle forces

1.2.4—Interpret the name and formula of compounds using IUPAC

1.2.5—Lewis Structure and Polarity

Problems

1. Describe covalent bonding.
2. Draw the Lewis Diagram for O2, I2, and N2. Indicate if they contain single, double, or triple bonds.
3. What type of elements are typically used in a covalent bond? Describe their electronegativity differences.
4. Using page 169, determine the type of bond from the electronegativity values for the following:
   1. N-Br b. Cl2 c. CaF2 d. CO
5. Explain why intermolecular forces are weaker than ionic, covalent or metallic bonds.
6. Explain why hydrogen bonds are stronger than dipole-dipole forces which are stronger than dispersion forces.
7. What is the relationship between bond energy and bond length?
8. Write the formula for the following:
   1. Carbon tetrachloride c. Dinitrogen pentoxide e. Sulfurous acid
   2. Chloric acid d. Nitric acid f. Hydrochloric acid
9. Write the names for the following formulas:
   1. PF5 c. H2SO4 e. H3PO4
   2. N2O3 d. HC2H3O2  f. HNO3
10. List 4 properties of covalent bonds and explain why they have these properties.
11. Complete the chart

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Molecule | (Want-have)/2 = # bonds | Lewis structure | Polarity | Geometry | IMF’s |
| NH3 |  |  |  |  |  |
| SO2 |  |  |  |  |  |
| CF4 |  |  |  |  |  |
| HBr |  |  |  |  |  |
| CO3-2 |  |  |  |  |  |

1. What is resonance and which of the molecules from #2 has resonance?
2. Analysis of a chemical indicates that it has a composition of 65.45%C, 5.45% H, and the rest is oxygen. The molar mass is found to be 165.0g/mol. Determine the empirical and molecular formulas.
3. Determine the empirical formula for the hydrate. Name the hydrate.

|  |  |
| --- | --- |
| Data for CaCl2 ∙ xH2O | |
| Mass of empty crucible | 25.40g |
| Mass of hydrate and crucible | 30.12g |
| Initial mass of hydrate |  |
| Anhydrous and crucible | 28.96g |
| Mass of anhydrous solid |  |
| Mass of water |  |

**Unit 6 Reactions and Stoichiometry**

Things to know: indicators of a reaction, characteristics of a chemical equation, word vs. formula reactions, symbols used in reactions, balancing reactions, five types of reactions, predicating products, mole ratio, stoichiometry, limiting reactant, excess reactant, percent yield

Objectives

2.2.2—Analyze the evidence of chemical change.

2.2.3—Analyze the law of conservation of matter and how it applies to chemical equations.

2.2.4—Analyze the stoichiometric relationships inherent in a chemical reaction.

Problems

1. List the four indicators of a chemical change.
2. Write a complete balanced reaction for each of the following:
   1. Magnesium oxide reacts with water.
   2. Silver reacts with chlorine.
   3. Zinc chlorate.
   4. Calcium carbonate.
   5. Lithium reacts with copper (II) chloride.
   6. Sodium reacts with hydrobromic acid.
   7. Chlorine reacts with potassium fluoride.
   8. Potassium sulfide reacts with zinc fluoride.
   9. Barium hydroxide reacts with nitric acid.
   10. Sodium chloride reacts with lead (II) nitrate.
   11. Glycerol (C3H8O3).
3. Use activity series to predict whether the following reactions take place.
   1. Cu + 2AgNO3🡪 Cu(NO3)2 + 2Ag Yes reaction occurs or No reaction
   2. 2KBr + I2🡪 2KI + Br2 Yes reaction occurs or No reaction
4. Use the solubility rules to determine the precipitate in each of the following reactions:
   1. NaCl + AgNO3🡪 NaNO3 + AgCl Precipitate: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
   2. 2NaOH + CuSO4🡪 Cu(OH)2 + Na2SO4 Precipitate: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
5. How many grams of zinc are needed to prepare 3.00 liters of H2 collected at STP?

Zn + H2SO4🡪 ZnSO4 + H2

1. How many grams of ammonium chloride are needed to make 0.100 mole of ammonia?

2NH4Cl + CaO🡪2NH3 + CaCl2 + H2O

1. How many particles of H2S can be made if you have 3.4moles of HCl? FeS + 2HCl 🡪 H2S + FeCl2
2. Chlorine is prepared by the reaction 2KMnO4 + 16HCl 🡪 2KCl + 2MnCl2 + 5Cl2 + 8H2O. How many moles of KMnO4 is needed to prepare 2.50 liters of Cl2 at STP?
3. Using the following reaction solve: 2FeCl3 + 3H2S 🡪 Fe2S3 + 6HCl. How many particles of HCl is produced when 90.0 g of FeCl3 reacts with excess H2S??
4. A 50.6 g sample of Mg(OH)2 is reacted with excess HCl according to the reaction below. How many grams of water are formed? Mg(OH)2 + 2 HCl --> MgCl2 + 2 H2O
5. Quicklime, CaO, can be prepared by roasting limestone, CaCO3, according to the chemical equation below. When 2.00 x 1023particles of CaCO3 are heated, how many liters of carbon dioxide are formed? CaCO3🡪CaO + CO2
6. Aluminum reacts with an aqueous solution containing excess copper (II) sulfate. If 4.27x1022 particles of copper II sulfate react with excess Al, what mass of copper is formed?

Al + CuSO4🡪 Cu + Al2(SO4)3

**Unit 7: Gas Laws**

Things to know*:* Kinetic-Molecular Theory of gases, converting pressures, converting temperatures, STP, Boyle’s Law, Charles’ Law, Gay-Lussac’s Law, combined gas law, volume-mass relationship, Dalton’s law of partial pressures, ideal gas law, effusion, diffusion, Graham’s law

Objective

2.1.5—Explain the relationships among pressure, temperature, volume, and quantity of gas, both qualitative and quantitative.

Problems

1. Identify characteristics of ideal gases.
2. What conditions allow gases to be the most soluble?
3. A helium balloon contains 125mL of gas at a pressure of 0.974atm. What will be the pressure if the volume is changed to 212mL?
4. A sample of gas in a closed container at a temperature of 102oC and a pressure of 3.2atm is heated. The new pressure was measured at 4.6atm. What was the final temperature in oC?
5. The volume of a gas is 23.5mL at 34.0oC and 1.25atm. What will the temperature be if the volume and pressure change to 18.7mL and 1.98atm?
6. At STP, what is the volume of 74.2g of nitrogen?.
7. Oxygen gas from the decomposition of KClO3 was collected over water. The total pressure was found to be 750.3torr. The water’s pressure was measured to be 23.4torr. What is the pressure of oxygen?
8. An engineer pumps 45.2g of carbon dioxide gas into a cylinder that has a capacity of 23.5L. What is the pressure inside the cylinder at 23.8oC?
9. A student collects 584mL of oxygen at a temperature of 25.8oC and a pressure of 683mmHg. How many grams of oxygen did the student collect?
10. If 3.2 mole of helium gas occupies 4.3L, what volume does 6.5 moles of helium occupy?

**Unit 8 Liquids and Solids**

Things to know*:* Kinetic Molecular Theory of liquids, Kinetic Molecular Theory of solids, types of solids, colligative properties, types of crystalline solids, phase changes, equilibrium during phase changes, order, endothermic or exothermic changes during a phase change, phase diagram, molar heat of vaporization, molar heat of fusion, heating and cooling curve

Objectives

2.1.1—Explain the energetic nature of phase changes.

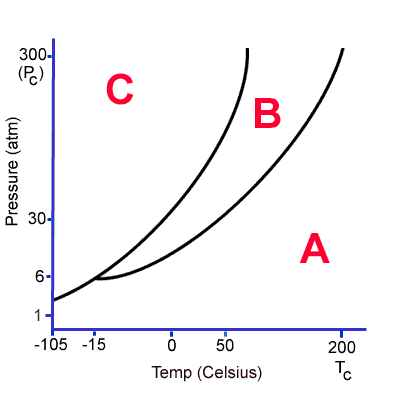
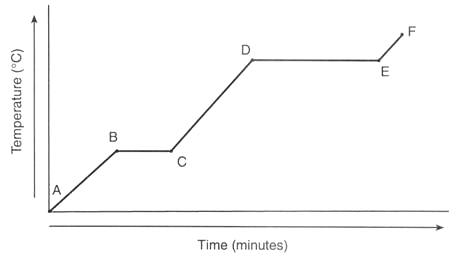
2.1.2—Explain heating and cooling curves.

2.1.3—Interpret the data presented in phase diagrams.

Problems

1. Explain physical equilibrium.
2. What two things does vapor pressure depend on?
3. Explain how the energy (kinetic and potential) of particles of a substance changes when heated, cooled, or changing phase.
4. What two conditions can change to cause a phase change?
5. Explain the difference between heat and temperature.
6. Fill out the chart

|  |  |  |
| --- | --- | --- |
| Phase Change | Name of Phase Change | Endothermic or Exothermic |
| Solid to Liquid |  |  |
| Liquid to Solid |  |  |
| Liquid to Gas |  |  |
| Gas to Liquid |  |  |
| Solid to gas |  |  |
| Gas to Solid |  |  |

1. Answer the following questions using the phase diagram.
   1. What does A represent?
   2. What does B represent?
   3. What does C represent?
   4. What is the P and T of the triple point?
   5. What phase is the substance in at 30atm and 200oC?
   6. At what T would the substance boil at 30atm?
   7. Draw an arrow representing sublimation.
2. What is the critical point?
3. Answer the following questions using the heating curve.
   1. What segment represents a solid being heated?
   2. What segment represents a liquid being heated?
   3. What segment represents a gas being heated?
   4. What segment represents melting?
   5. What segment represents condensation?

**Unit 9 Solutions**

Things to know*:* types of mixtures, suspensions, colloids, electrolytes vs. non-electrolytes, factors affecting rate of dissolving, saturated vs. unsaturated vs. supersaturated, “like dissolves like”, gas solubility, solubility curve, molarity, dilution, *molality*

Objectives

3.2.3—Infer the quantitative nature of a solution.

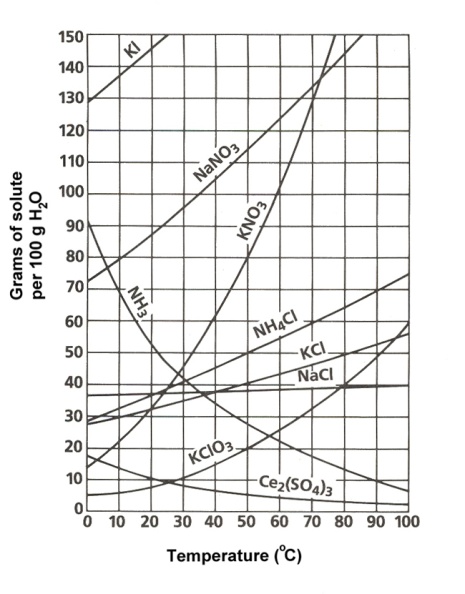
3.2.4—Summarize the properties of solutions.

3.2.5—Interpret solubility diagrams.

3.2.6—Explain the solution process.

2.2.3—Solubility Rules.

Problem

1. What is the molarity of a solution containing 2.0 grams of NaCl in 30.mL of water?
2. How many grams must be dissolved to make 2.5L of a 5.0M solution of HCl?
3. How many liters of 3.00M HF solution must be used to dilute 5.0L of 6.00M HF?
4. Define solution.
5. What happens to the following as more solute is added?
   1. vapor pressure b. boiling point c. freezing point d. osmotic pressure
6. Which of the following would have the highest boiling point?
   1. 1.2M NaCl b. 1.5M CaCl2 c. 2.0M CO2 d. 1.0M C12H22O11
7. Define the following as electrolyte or nonelectrolyte:
   1. **HCl b. CO2 c. LiBr d. NaOH e. ethyl alcohol
8. Use the solubility curve to answer the following questions
   1. How many grams of sodium nitrate can be dissolved in 100g of

water at 40oC?

* 1. At what temperature can 22 grams of potassium chlorate be

dissolved in 100g of water?

* 1. What salt is the most soluble at 70oC?
  2. What salt is the least soluble at 30oC?
  3. If 100g of potassium iodide is dissolved at 10oC, is the solution

saturated, supersaturated, or unsaturated?

1. Develop a conceptual model for the solution process with a cause and effect relationship involving forces of attraction between solute and solvent particles.
2. Describe the energetics of the solution process as it occurs and the overall process as exothermic or endothermic.
3. Label each of the following as solid (s) or aqueous (aq).
   1. AgBr b. NaC2H3O2 c. Ca3(PO4)2 d. Al2S3

**Unit 10 Chemical Equilibrium**

Things to know: equilibrium, dynamic equilibrium, equilibrium constant, K>1, K<1, homogenous equilibrium, heterogeneous equilibrium, determining equilibrium constant, finding equilibrium concentration, Le Châtelier’s principle

Objectives

3.1.2—Explain the conditions of a system at equilibrium.

3.1.3—Infer the shift in equilibrium when a stress is applied to a chemical reaction.

Problems

1. Define chemical equilibrium.
2. If K is less than 1, what does that tell us about equilibrium for that reaction?
3. If K is greater than 1, what does that tell us about equilibrium for that reaction?
4. Write the equilibrium constant for the following reactions:
   1. 2N2H4(g) + 2NO2(g) ↔ 3N2(g) + 4H2O(g)
   2. I2(g)↔ 2I(g)
   3. Fe3O2(s) + 4H2(g) ↔ 3Fe(s) + 4H2O(l)
5. When solid ammonium chloride is put in a reaction vessel at 323K, the equilibrium concentrations of both ammonia and hydrogen chloride are found to be 0.0660M. Calculate Keq. NH4Cl(s) ↔ NH3(g) + HCl(g)
6. Calculate K for the following equilibrium when [SO3] = 0.0160M, [SO2] = 0.00560M, and [O2] = 0.00210M. 2SO3(g) ↔ 2SO2(g) + O2(g)
7. How would increasing the volume of the reaction vessel affect these equilibria?
   1. NH4Cl(s) ↔ NH3(g) + HCl(g)
   2. N2(g) + O2(g) ↔ 2NO(g)
8. Use Le Châtelier’s principle to predict how each of the following changes would affect this equilibrium. H2(g) + CO2(g) ↔ H2O(g) + CO(g)
   1. Adding H2O(g) to the system
   2. removing CO (g) from the system
   3. adding H2(g) to the system
   4. adding something to the system to absorb CO2(g)
9. How would simultaneously decreasing the temperature of the system affect these equilibria?
   1. Heat + CaCO3(s) ↔ CaO(s) + CO2(g)
   2. 4NH3(g) + 5O2(g) ↔ 4NO(g) + 6H2O(g) + heat

**Unit 11 Acid Base**

Things to know: net ionic equations, strong vs. weak electrolytes, properties of acids and bases, Arrhenius acids and bases, labeling acid/bases and conjugate acid/base, Brønstead-Lowry acids and bases, monoprotic vs. polyprotic, amphoteric compound, neutralization reactions, calculating pH, pOH, [H+], [OH-], acidic vs. basic based on pH, titration, indicators, endpoint vs. equivalence point, calculations involving titrations

Objectives

3.2.1—Classify substance using the [H3O]+ and [OH]-

3.2.2—Summarize the properties of acids and bases.

3.2.3—Infer the quantitative nature of a solution.

2.2.3—Analyze the law of conservation of matter and how it applies to chemical equations.

Problem

1. Label the acid, base, conjugate acid, and conjugate base in the following reactions:
   1. HNO3 + H2O ↔ H3O+ + NO3- b. NH3 + H2O ↔ NH4+ + OH-
2. Which of the following is the most concentrated acid: 1.0x10-5M HCl or 1.0x10-2M HF
3. Which of the following is a strong acid: 1.0x10-5M HCl or 1.0x10-2M HF
4. What is the difference between a strong acid or base and a weak acid or base?
5. What color would phenolphthalein turn if the solution had a pH of 8.5?
6. What color would the litmus paper turn if the solution had a pH of 3.4?
7. Calculate the pH at 298K of a solution having the hydrogen ion concentration of 1.0x10-11M.
8. Calculate the pOH and the pH at 298K of a solution having [OH-] = 1.3x10-5M.
9. Label each of the following Acid, Base, Both, or Neutral.
   1. Can turn litmus paper blue.
   2. Reacts with certain metals.
   3. Contains more hydronium ions than hydroxide ions.
   4. Has a pH of 7.
   5. Feels slippery.
   6. Keeps phenolphthalein clear.
   7. Tastes bitter.
   8. Tastes sour.
   9. Is an electrolyte.
   10. Can turn litmus paper pink.
   11. Occurs when a strong acid reacts with a strong base.
   12. Turns phenolphthalein pink.
   13. Contains more hydroxide ions than hydronium ions.
   14. Has a pH of 2.
10. In a titration, 33.21mL of 0.3020M sodium hydroxide solution is required to exactly neutralize 20.00mL hydrofluoric acid solution. What is the molarity of the hydrofluoric acid solution?
11. Write the net ionic equation for copper (II) chloride reacting with ammonium phosphate. List any spectator ions.
12. Write the net ionic equation for the reaction between barium hydroxide and hydro bromic acid. List any spectators.

**Unit 12 Energy and Kinetics**

Things to know: collision theory, thermochemistry, heat vs. temperature, calorimeter, heat capacity and specific heat, heat of vaporization, heat of fusion, enthalpy, potential energy diagram, activation energy, *heats of formation*, *Hess’ law*, *entropy*, kinetics, reaction rate, rate-influencing factors, *average rate*, *ratelaw*, *rate-determining step* (specific heats, ΔHv= 2260J/g, ΔHf=334J/g are on front of reference table)

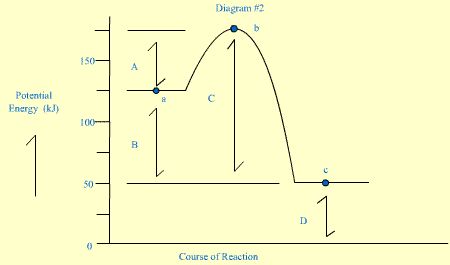
Objectives

3.1.1—Explain the factors that affect the rate of a reaction.

2.2.1—Explain the energy content of a chemical reaction.

2.1.4—Infer simple calorimetric calculations based on the concepts of heat lost equals heat gained and specific heat.

Problems

1. What three things must happen for a reaction to take place?
2. What is happening to the bonds when reactions are taking place?
3. What happens to the reaction rate when temperature is increased? Explain.
4. What happens to the reaction rate when concentration is increased? Explain.
5. What happens to the reaction rate when pressure is increased? Explain.
6. What happens to the reaction rate when surface area is increased? Explain.
7. What happens to the reaction rate when a catalyst is added? Explain.
8. What is the sign for ΔH when the reaction is exothermic? Endothermic?
9. The temperature of a 94.8 grams sample of material increases from 37oC to 79oC when it absorbs 3584J of heat. What is the specific heat of this material? What is this material? (p492)
10. What mass of liquid water at room temperature can be raised to its boiling point with addition of 24kJ of heat energy?
11. How much heat is absorbed when 5.980 grams of water is melted?
12. Label the potential energy diagram below.

Is this an endothermic or exothermic reaction?