**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:
   1. Formation of a precipitate
   2. Formation of a gas
   3. Change in color
   4. Change in energy
2. Define precipitate. How can you determine if something is a precipitate or not?

New solid formed during a reaction…is always on the product side (right side)…can determine a precipitate by using the solubility rules (insoluble = precipitate)

1. What would happen if a burning splint was exposed to:
   1. Oxygen the flame would be brighter and bigger
   2. Hydrogen the flamewould make a popping sound…sometimes called barking
   3. Carbon dioxidethe flame would go out
2. What would happen if carbon dioxide was bubbled through lime water?

The water would look cloudy

1. What piece of safety equipment must be used in all laboratory experiments?

Goggles

1. When using a Bunsen burner what precautions should be made?

Long hair and clothes should be pulled back and there should be no flammable materials near by

1. Exothermic reaction releases energy, ∆H is neg and energy is shown on the product (right)side.
2. Endothermic reaction absorbs energy, ∆H is + and energy is shown on the reactant (left)side.
3. Label each of the following as a physical or chemical property and give a brief reason
   1. Flammability c. Can neutralize a base

CP, cannot get substance back/change in color CP, cannot get substance back

* 1. Density d. Boiling Point

PP, description…doesn’t change substance PP, phase change…still same substance

1. 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.

PC, doesn’t change substance PP, diluted…same substance…no color change

* 1. Food is digested in the stomach. d. Water is heated and changed into steam.

CP, can’t get substance back, gas formed PP, phase change….still same substance

**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.

P+: positive, nucleus, 1amu….no: neutral, nucleus, 1amu…e-: negative, electron cloud, 0amu

1. Isotopes differ in the number of neutrons not protons(which always must stay the same as they indicate the atomic number!)

Ex. AZX Ex. F-20

Z means atomic number (p+) Atomic Number 9 (from periodic table)

A means mass number (p+ + no) Mass Number 20

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

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Isotope | Electrons | Protons | Mass Number | Atomic Number | Neutrons |
| 238U | 92 | 92 | 238 | 92 | 146 |
| 16O2- | 10 | 8 | 16 | 8 | 8 |
| 23Na+1 | 12 | 11 | 23 | 11 | 12 |

1. What is the difference between average atomic mass, isotopic mass, and mass number?

Average atomic mass: average of all isotopes…isotopic mass/mass number: one specific isotope

1. Balancing Nuclear Reactions (remember Law of Conservation of Mass) and label fission or fusion
   1. Ar🡪3717Cl + 11H ­­­­­­­­­­­­­fission b. 6430Zn + H 🡪6429Cu + 10nfusion (1 isotope at end)
2. Give the symbol for the following
   1. Alphaα, He b. Beta β, e c. Gammaγ, energy
3. Which type of decay listed above is the strongest? The weakest? Explain why.

Gamma is the strongest because it is just energy, alpha is the weakest because it is the heaviest

1. Decay is a random event, independent of other energy influences. Cannot stop it from happening
2. Balance the following nuclear reactions and label alpha decay or beta decay
   1. 220Rn 🡪4He + Po alpha b. 216Po 🡪0-1e + At beta
3. 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?

40.0g 🡪 20.0g 🡪 10.0g 🡪 5.00g 🡪 2.50g 🡪 1.25g

12.3yrs + 12.3 yrs + 12.3yrs + 12.3 yrs + 12.3 yrs = 61.5 years

1. Fe-61 has a half-life of 6.00min. Of a 100.0g sample, how much will remain after 18.0min?

18.0 minutes = 3 half-lives 100.0g 🡪 50.0g 🡪 25.0g 🡪12.5g

6.00 minutes

1. After 20.0 days, a 120kg sample of Bi-210 decays down to just 7.5kg. What is its half-life?

120kg 🡪 60kg 🡪 30kg 🡪 15kg 🡪 7.5kg 4 half lives

20.0 days = 5.00 days

4 half lives

1. What is the molar mass of CO2?

12.0g/mol + 2(16.0g/mol) = 44.0g/mol

1. What is the molar mass of (NH4)2S • 3H2O?

2(14.0g/mol) + 8(1.0g/mol) + 32.1g/mol + 3(18.0g/mol) = 122.1g/mol

1. How many grams are in 3.0 moles of H2SO4?

3.0molH2SO4 (98.1g H2SO4) = 290g H2SO4 (this is with correct number of sig figs)

(1mol H2SO4)

1. How many molecules are in 64 grams of O2?

64g O2 (6.022x1023moleucles O2) = 1.2x1024molecules O2

( 32 g O2)

1. How many moles are in 84.2 grams of CO2?

84.2g CO2 (1 mole CO2) = 1.91 moles CO2

(44g CO2)

1. How many moles are in 3.04x1023 molecules of H2?

3.04x1023molecules H2( 1 mole H2) = 0.505 mol H2

(6.022x1023molecules H2)

1. How many grams are in 4.59x1025 particles of NaCl?

4.59x1025particles NaCl ( 58.5g NaCl) = 4460g NaCl

(6.022x1023particles 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

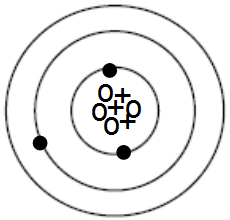
1. According to Bohr’s model of the atom, where is the only place electrons can be located?

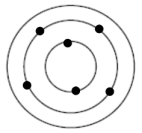
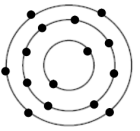
In circular orbits

1. Describe the electron cloud.

A 3d region of space where electrons are most likely to be found…not circling the nucleus but contained in different probability shapes (based on Schrödinger’s equation)

1. Use the diagram below to draw the 7Li isotope.



1. Label each diagram with the element (***name and symbol***) they represent.
   1. carbon, C
   2. phosphorus, P
2. Define quanta. Contains a specific amount of energy
3. Describe the difference between the ground state and an excited state. How do electrons move from one to the other?

Ground state is the lowest energy level, closest to the nucleus…Excited state is higher energy…electrons must gain energy to move to higher energy level and they lose energy to move to a lower energy level

1. Define photon.

A quanta of energy emitted when an electron falls from an excited state to lower energy level

1. What is the relationship between(longer wavelength, lower energy, lower frequency)
   1. Wavelength and frequency inverse b. energy and frequency direct
2. 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?

656nm, visible, red color

1. 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?

1094nm, IR

1. Describe the wave/particle duality of electrons.

Electrons can act like a wave (give off photons) and a particle (have mass and take up space)

1. Niels Bohr produced a model of the hydrogen atom based on experimental observations. This model indicated that:
   1. An electron circles the nucleus only in fixed energy ranges called orbits.
   2. An electron can neither gain nor lose energy within this orbital, but could move up or down to another orbit.
   3. The lowest energy orbit is closest to the nucleus.
2. Define groups.

Columns, vertical, up/down, also called families

1. Name three things main group elements in the same family have in common.

Number of valence electrons

Oxidation number (charge)

Properties

1. Reactivity increases down a group for metals and decreases for nonmetals.
2. Define periods.

Rows, horizontal, left/right

1. Where are the following elements located on the periodic table:
   1. Metals b. Nonmetals c. Metalloids

Left of stairs Right of stairs along stairs (not Al)

1. Label the following elements as a metal, nonmetal, or metalloid
   1. Oxygen b. lead c. silicon d. magnesium e. boron f. neon

Nonmetal metal metalloid metal metalloid nonmetal

1. Give the location of the following families on the periodic table
   1. Representative 1-2, 13-18, A
   2. Alkali metalsGroup 1
   3. Alkaline earth metals Group 2
   4. Halogens Group 17
   5. Noble gases Group 18
   6. Transition elements3-12, B’s
2. Define atomic radius.

Size of atom

1. Define ionic radius.

Size of ion, atom with a charge

1. What is the group and period trend for atomic radius?

Increase down a group and decrease across a period (right to left)

1. Put the following in order by decreasing atomic radius: Al, Na, S, K and explain why.

K, Na, Al, S…K has more energy levels since it is in a higher row…as you add more protons to the same energy level the electrons are attracted more to the nucleus so the atom becomes smaller

1. Which of the following has the most metallic character: Al, Na, S, or K?

K…it is the largest atom so it will lose electrons more easily

1. Write the electron configuration for Li and F. Explain which will lose electrons and which will gain electrons to become stable.

Li: 1s22s1 F: 1s22s22p5 Li will lose electrons because it has 1 valence electron and it is easier to lose 1 than gain 7. F will gain 1 electron because it has 7 valence electrons and it is easier to gain 1 than lose 7

1. The more metallic an element is the lower the ionization energy, lower the electron affinity, and the lower the electronegativity.
2. Name the four sublevels and the area it represents on the periodic table.

s: groups 1 and 2 and He p: groups 13-18 d: groups 3-10 f: bottom 2 rows

1. Write the orbital notation for the following:
   1. Sulfur

↑↓ ↓↑ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑ ↑

1s 2s 2p 3s 3p

* 1. Nickel

↑↓ ↓↑ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑ ↑

1s 2s 2p 3s 3p 4s 3d

1. Write the electron configuration for the following elements
   1. Boron

1s22s22p3

* 1. Copper

1s22s22p63s23p64s23d9

1. Identify the element represented by the configurations below
   1. 1s22s22p4 b. [Ar]4s23d104p3

Oxygen Arsenic

1. Determine the number of valence electrons in the following configurations
   1. 1s22s22p63s23p64s23d104p5 b. [Kr]5s1

2 + 5 = 7 valence electrons 1 valence electron

1. List the number of valence electrons for each of the following
   1. Sodium b. Nitrogen c. Bromine

Group 1 = 1 Group 15 = 5 Group 17 = 7

1. Using the configurations below, determine the number of electrons lost or gained and the oxidation number it will form
   1. 1s22s22p63s23p64s23d104p65s2 b. 1s22s22p63s23p4

2 valence electrons = +2 2 + 4 = 6 valence electrons = -2

1. Define ionization energy.

Energy needed to remove an electron from an atom in the gaseous state

1. What is the general trend for ionization energy? Explain the reasoning.

Decrease down and increase across (left to right)…the smaller the atom the more tightly held the electron is to the nucleus making the electron harder to remove…the larger the atom the electron is not held as tightly so easier to remove

1. Put the following elements in order of decreasing ionization energy: Rb, Al, S, Mg

S, Al, Mg, Rb (for larger atoms the easier it is to remove the electron)

1. Define electronegativity.

Energy needed to gain an electron when the atom is in the gaseous state

1. What is the general trend for electronegativity? Explain the reasoning.

Decrease down and increase across (left to right)…the smaller the atom the more tightly held the electron is to the nucleus making an extra electron easier to be attracted…the larger the atom the electron is not held as tightly so an extra electron has a harder time being attracted

1. Put the following elements in order of increasing electronegativity energy: F, B, N, Li

Li, B, N, F

**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.

Metal atoms where the valence electrons for a sea of electrons (delocalized/mobile)

1. How are ions formed? Which arrangements are stable?

Electrons are lost or gained to achieve a stable state: full or half filled s and p sublevels

1. Explain difference between cation and anion.

Cation is pos charged ion (electrons lost) and Anion is a neg charged ion (electrons gained)

1. 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

+ 1 +2 +3 -3 -2 -1

1. What types of elements form ionic compounds? Explain their electronegativity differences.

Metals and nonmetals…cations and anions…electronegativity difference above 1.7

1. Predict the chemical formulas of compounds using Lewis structures.
   1. Potassium and Sulfur b. Magnesium and Oxygen  
2. Give four properties of ionic compounds and explain why they have these properties.

High melting point, high boiling point, brittle, and high electrical conductivity either in molten state or in aqueous solution…these properties exist because valence electrons are being transfered

1. Give six properties of metallic compounds and explain why they have these properties.

High melting point, high boiling point, high conductivity, malleability, ductility, and luster…these properties exist because valence electrons are creating a sea of electrons (delocalized/mobile)

1. Write the formula for the following:
   1. Magnesium fluorideMgF2 d. Calcium nitrideCa3N2
   2. Sodium carbonateNa2(CO3) e. Ammonium phosphate(NH4)3(PO4)
   3. Copper (III) bromideCuBr3 f. Tin (IV) oxideSnO2
2. Write the names for the following formulas:
   1. FeS iron (II) sulfide d. VBr3vanadium (III) bromide
   2. NH4NO2 ammonium nitritee. Ba(OH)2barium hydroxide
   3. Al2S3 aluminum sulfidef. (NH4)3Pammonium phosphide
3. Calculate the percent by mass of water in lithium chromate dihydrate.

Li2CrO4∙ 2H2O 2(6.9g/mol) + 52.0g/mol + 4(16.0g/mol) + 2(18.0g/mol) = 165.8g/mol

2(18.0g/mol) x 100 = 21.7%H2O

165.8g/mol

**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.

Sharing of electrons to achieve a stable octet

1. Draw the Lewis Diagram for O2, I2, and N2. Indicate if they contain single, double, or triple bonds.O2 contains double bond, I2 contains a single bond, N2 contain a triple bond



1. What type of elements are typically used in a covalent bond? Describe their electronegativity differences.

Nonmetal elements, electronegativity values are closer

1. Using page 169, determine the type of bond from the electronegativity values for the following:
   1. N-Br b. Cl2 c. CaF2 d. CO

3-2.8=0.2 nonpolar 3-3=0 nonpolar 4-1=3 ionic 3.5-2.5=1 polar

1. Explain why intermolecular forces are weaker than ionic, covalent or metallic bonds.

IMF are weaker than ionic, covalent, or metallic because the electrons are being shared instead of completely transferred or freely moving

1. Explain why hydrogen bonds are stronger than dipole-dipole forces, which are stronger than dispersion forces.

Hydrogen bonding is stronger than dipole-dipole which is stronger than dispersion forces because there is a stronger dipole moment (electrons are more polarizable…they are hanging out with the more electronegative atom more often)

1. What is the relationship between bond energy and bond length?

The more bonds between two atoms the stronger the bond (requires more energy to break bond)…triple bond is the shortest and strongest

1. Write the formula for the following:
   1. Carbon tetrachloride CCl4c. DinitrogenpentoxideN2O5 e. Sulfurous acidH2SO3
   2. Chloric acidHClO3 d. Nitric acid HNO3  f. Hydrochloric acid HCl
2. Write the names for the following formulas:
   1. PF5phosphorus pentafluoride c. H2SO4sulfuric acid e. H3PO4phosphoric acid
   2. N2O3dinitrogen trioxided. HC2H3O2acetic acidf. HNO3 nitric acid
3. List 4 properties of covalent bonds and explain why they have these properties.

Low melting point, low boiling point, poor electrical conductivity, polar nature…these properties exist because the valence electrons are being shared

1. Complete the chart

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Molecule | (Want-have)/2 = # bonds | Lewis structure | Polarity | Geometry | IMF’s |
| NH3 | (14 – 8) = 3 bonds  2 |  | Polar | Trigonal pyramidal | Hydrogen bonding |
| SO2 | (24-18) = 3 bonds  2 |  | Polar | Bent | Dipole-dipole |
| CF4 | (40-32) = 4 bonds  2 |  | Nonpolar | Tetrahedral | London dispersion |
| HBr | (10-8) = 2 bonds  2 |  | Polar | Linear | Dipole-dipole |
| CO3-2 | (32-24) = 4 bonds  2 |  | Nonpolar | Trigonal Planar | London Dispersion |

1. What is resonance and which of the molecules from #2 has resonance?

SO2 and CO3-2

1. 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.

65.45gC (1 mole C) = 5.45mol C = 3 carbon

(12.0g C ) 1.82 mol Empirical Formula: C3H3O

5.45gH (1mole H) = 5.45molH = 3 hydrogen Molecular Formula: C9H9O3

( 1.0 g H ) 1.82 mol

100%-65.45%C-5.45%H=29.1%O 29.1gO (1mol O) = 1.82molO = 1 oxygen 165.0g/mol = 3

(16.0gO) 1.82 mol 55.0g/mol

1. 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 | (30.12-25.40g) = 4.72g |
| Anhydrous and crucible | 28.96g |
| Mass of anhydrous solid | (28.96g-25.40g) = 3.56g |
| Mass of water | (4.72g-3.56g) = 1.16g |

3.56gCaCl2 (1mol CaCl­2) = 0.0321molCaCl2 = 1 CaCl2

( 111g CaCl2) 0.0321mol

1.16gH2O (1molH2O) = 0.0644mol H2O = 2 H2O

(18.0gH2O) 0.0321mol

Empirical Formula: CaCl2 ∙ 2H2O

Name: calcium chloride dihydrate

**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.

Formation of precipitate, formation of gas, change in color, change in energy

1. Write a complete balanced reaction for each of the following:
   1. Magnesium oxide reacts with water.

MgO + H2O 🡪Mg(OH)2

* 1. Silver reacts with chlorine.

2Ag + Cl2🡪 2AgCl

* 1. Zinc chlorate.

Zn(ClO3)2🡪 ZnCl2 + 3O2

* 1. Calcium carbonate.

CaCO3🡪CaO + CO2

* 1. Lithium reacts with copper (II) chloride

2Li + CuCl2🡪 2LiCl + Cu

* 1. Sodium reacts with hydrobromic acid.

2Na + 2HBr 🡪 2NaBr + H2

* 1. Chlorine reacts with potassium fluoride.

Cl2 + KF 🡪 no reaction because Cl2 is lower on activity series than F2

* 1. Potassium sulfide reacts with zinc fluoride.

K2S(aq) + ZnF2(aq) 🡪 2KF(aq) + ZnS(s) (ZnS would be the precipitate formed)

* 1. Barium hydroxide reacts with nitric acid.

Ba(OH)2(aq) +2HNO3(aq) 🡪 Ba(NO3)2(aq) + 2H2O(l) (double replacement of acid/base)

* 1. Sodium chloride reacts with lead (II) nitrate.

2NaCl(aq) + Pb(NO3)2(aq) 🡪 2NaNO3(aq) + PbCl2(s) (PbCl2 would be the precipitate)

* 1. Glycerol (C3H8O3).

2C3H8O3 + 7O2🡪6CO2 + 8H2O

1. Use activity series to predict whether the following reactions take place.
   1. Cu + 2AgNO3🡪 Cu(NO3)2 + 2Ag Yes reaction occurs…Cu is higher than Ag
   2. 2KBr + I2🡪 2KI + Br2 No reaction…I2 is lower than Br2
2. Use the solubility rules to determine the precipitate in each of the following reactions:
   1. NaCl + AgNO3🡪 NaNO3(aq)+ AgCl(s) Precipitate: AgCl is the solid formed
   2. 2NaOH + CuSO4🡪 Cu(OH)2(s) + Na2SO4(aq)Precipitate: Cu(OH)2 is the solid formed
3. How many grams of zinc are needed to prepare 3.00 liters of H2 collected at STP?

Zn + H2SO4🡪 ZnSO4 + H2

3.00L H2 (1mol H2) (1mol Zn) (65.4g Zn) = 8.76g Zn

(22.4L H2)(1mol H2) (1mol Zn)

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

2NH4Cl + CaO🡪2NH3 + CaCl2 + H2O

0.100mol NH3 (2mol NH4Cl) (53.5gNH4Cl) = 5.35g NH4Cl

(2mol NH3 ) (1mol NH4Cl)

1. How many particles of H2S can be made if you have 3.4moles of HCl? FeS + 2HCl 🡪 H2S + FeCl2

3.4mol HCl (1mol H2S) (6.022x1023paricles H2S) = 1.0x1024 particles H2S

(2mol HCl) ( 1mol H2S )

1. 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?

2.50L Cl2 (1mol Cl2) (2mol KMnO4) = 0.0446mol KMnO4

(22.4LCl2) (5 mol Cl2 )

1. 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??

90.0g FeCl3 (1mol FeCl3) (6mol HCl) (6.022x1023particles HCl) = 1.00x1024paricles HCl

(162.4gFeCl3) (2molFeCl3) (1 molHCl )

1. A 50.6 g sample of Mg(OH)2 is reacted with excessHCl according to the reaction below. How many grams of water are formed? Mg(OH)2 + 2 HCl --> MgCl2 + 2 H2O

50.6g Mg(OH)2 (1mol Mg(OH)2) ( 2 mol H2O ) (18.0g H2O) = 31.2g H2O

(58.3g Mg(OH)2) (1molMg(OH)2) (1mol H2O)

1. 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

2.00x1023particles CaCO3 ( 1 mole CaCO3 ) ( 1mol CO2 ) (22.4L CO2) = 7.44L CO2

(6.022x1023paricles CaCO3) (1molCaCO3) (1mol CO2)

1. 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?

2Al + 3CuSO4🡪3Cu + Al2(SO4)3

4.27x1022particles CuSO4 ( 1 mole CuSO4 ) ( 3 mole Cu ) (63.6g Cu) = 4.51g Cu

(6.022x1023paricles CuSO4) (3molCuSO4) (1mol Cu)

**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.

Problems

1. Identify characteristics of ideal gases.

Volume the gas particles take up is zero compared to distance between them, collisions are elastic, gas particles are in constant rapid/random motion, gas particles have no repulsion or attractive forces between them, temperature is directly related to kinetic energy

1. What conditions allow gases to be the most soluble?

Gases are most soluble at low temperature and high pressure

1. 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?

(125mL x 0.974atm) = (212mL)X 121.75mL atm = (212mL)X 121.75mL atm = 0.574atm

121mL

1. 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?

273 + 102o = 375K

(3.2atm) = (4.6atm) (3.2atm)x = 1725atm K x = 1725atm K = 539K -273 = 266oC

(375K ) x 3.2atm

1. 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?

273 + 34.0oC = 307K

(23.5mL ∙ 1.25atm) = (18.7mL ∙ 1.98atm) (29.4mL atm)x = 11367mL atm K

(307K) x x = 11367mL atmK = 387K

29.4mL atm

1. At STP, what is the volume of 74.2g of nitrogen?

74.2g N2 (22.4L N2) = 59.4L N2

(28.0gN2)

1. 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?

750.3torr = 23.4torr + x 750.3torr – 23.4torr = x x = 726.9torr

1. 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 23.8 oC?

45.2g CO2 (1mol CO2) = 1.03 mol CO2 273 + 23.8oC = 296.8K

(44g CO2 )

X(23.5L) = (1.03mol ∙ 0.0821 ∙ 296.8K) X(23.5L) = 25.1L atm X = 25.1mL atm = 1.07atm

23.5L

1. 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?

584mL = 0.584L 683mmHg (1atm/760mmHg) = 0.899atm 273 + 25.8oC = 298.8K

(0.584L ∙ 0.899atm)=x (0.0821 ∙ 298.8K) 0.525L atm = x(24.5) 0.525L atm =

(24.5)

x = 0.024molO2 (32gO2/1molO2)=**0.686gO2**

1. If 3.2 mole of helium gas occupies 4.3L, what volume does 6.5 moles of helium occupy?

(3.2mol) = (6.5mol) (3.2mol)x = 1.5mol L x = (1.5mol L) = 0.47L

4.3L X 3.2 mol

**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.

Two opposite phase changes happening at the same rate

1. What two things does vapor pressure depend on?

Vapor pressure depends on concentration and temperature

1. Explain how the energy (kinetic and potential) of particles of a substance changes when heated, cooled, or changing phase.

Kinetic energy will increase when temperature increases, decrease when temperature decreases and remains the same during a phase change…Potential energy remains same when temperature is changing but changes during a phase change.

1. What two conditions can change to cause a phase change?

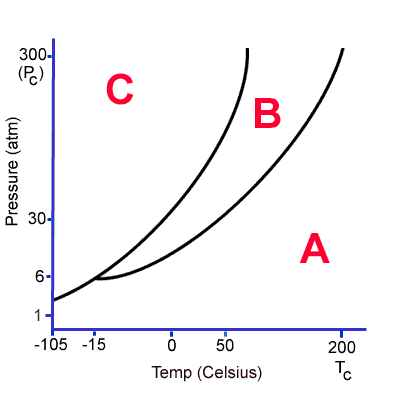
A change to temperature and/or pressure can cause a phase change

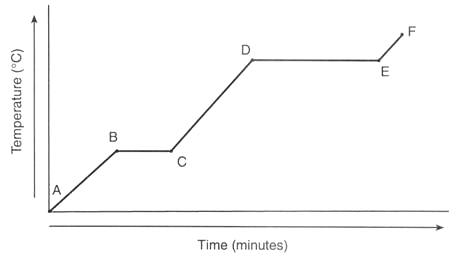
1. Explain the difference between heat and temperature.

Heat is the total kinetic energy but temperature is the average kinetic energy

1. Fill out the chart

|  |  |  |  |
| --- | --- | --- | --- |
| Phase Change | Name of Phase Change | Endothermic or Exothermic | |
| Solid to Liquid | Melting | Endothermic | |
| Liquid to Solid | Freezing | Exothermic | |
| Liquid to Gas | Boil/evaporation/vaporization | Endothermic | |
| Gas to Liquid | Condensation | Exothermic |
| Solid to gas | Sublimation | Endothermic |
| Gas to Solid | Deposition | Exothermic |

1. Answer the following questions using the phase diagram.
   1. What does A represent? Gas phase
   2. What does B represent? Liquid phase
   3. What does C represent? Solid phase
   4. What is the pressure and temp of the triple point?6atm/-15oC
   5. What phase is the substance in at 30atm and 200oC?gas
   6. At what temperature would the substance boil at 30atm?50-60oC
   7. Draw an arrow representing sublimation. Arrow would cross line from C towards A without going through the B region of the graph
2. What is the critical point?

The point at which a substance can no longer be liquefied

1. Answer the following questions using the heating curve.
   1. What segment represents a solid being heated? A-B
   2. What segment represents a liquid being heated? C-D
   3. What segment represents a gas being heated? E-F
   4. What segment represents melting? B-C
   5. What segment represents condensation? E-D

**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?

2.0g NaCl (1 molNaCl) = 0.0342mol NaCl 30.mL = 0.030L X = 0.0342mol= 1.1M

(58.5g NaCl) 0.030L

1. How many grams must be dissolved to make 2.5L of a 5.0M solution of HCl?

5.0M = x (5.0M ∙ 2.5L) = x x = 12.5mol HCl (36.5g HCl) = 456g HCl

2.5L (1mol HCl)

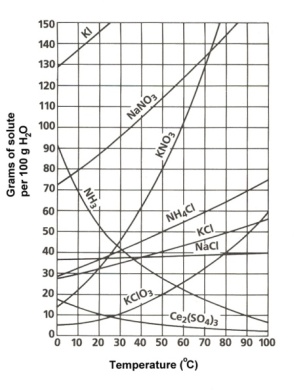
1. How many liters of 3.00M HF solution must be used to dilute 5.0L of 6.00M HF?

(6.00M ∙ 5.0L) = (3.00M)x (30.M L) = x x = 10.L

(3.00M)

1. Define solution.

Homogenous mixture…cannot see the different substances that make up the mixture because all in same phase

1. What happens to the following as more solute is added?
   1. vapor pressure↓ b. boiling point ↑ c. freezing point ↓ d. osmotic pressure↑
2. Which of the following would have the highest boiling point?
   1. 1.2M NaClb. 1.5M CaCl2 most amount of ions c. 2.0M CO2 d. 1.0M C12H22O11
3. Define the following as electrolyte or nonelectrolyte:
   1. HCl electro b. CO2nonelectroc. LiBr electro d. NaOH electro e. ethyl alcohol nonelectro
4. **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?100-110g

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

dissolved in 100g of water?50-60oC

* 1. What salt is the most soluble at 70oC?NaNO3
  2. What salt is the least soluble at 30oC?Ce2(SO4)3
  3. If 100g of potassium iodide is dissolved at 10oC, is the solution

saturated, supersaturated, or unsaturated? Less than amount

can dissolve

1. Develop a conceptual model for the solution process with a cause and effect relationship involving forces of attraction between solute and solvent particles.

The partial positive side of water molecule (H ends) will attract the negative ion and help pull it away from the solute. The partial negative side of water (O end) will attract the positive ion and help pull it way from the solute.

1. Describe the energetics of the solution process as it occurs and the overall process as exothermic or endothermic. Bonds break (reactants) and then bonds form (products) if more energy is used to break the bonds then released when bonds are formed it is endothermic…if less energy is used to break the bonds then released when bonds are formed it is exothermic.
2. Label each of the following as solid (s) or aqueous (aq).
   1. AgBr(s) b. NaC2H3O2(aq)c. Ca3(PO4)2(s) d. Al2S3(s)

**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.

The forward reaction is happening at equal rate as the reverse reaction.

1. If K is less than 1, what does that tell us about equilibrium for that reaction?

At equilibrium, the concentration of the reactants is larger than the products…reverse reaction favored

1. If K is greater than 1, what does that tell us about equilibrium for that reaction?

At equilibrium, the concentration of the products is larger than the reactants…forward reaction favored

1. Write the equilibrium constant for the following reactions:
   1. 2N2H4(g) + 2NO2(g) ↔ 3N2(g) + 4H2O(g)

K = [N2]3 [H2O]4 homogenous equilibrium since all gases

[N2H4]2[NO2]2

* 1. I2(g)↔ 2I(g)

K = [I]2 homogenous equilibrium since all gases

[I2]

* 1. Fe3O2(s) + 4H2(g) ↔ 3Fe(s) + 4H2O(l)

K = 1 heterogeneous equilibrium since different states

[H2]4 (no liquids or solids are included in equilibrium expression)

1. 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)(remember no solids)

K = [NH3] [HCl] K = 0.0660M ∙ 0.0660M = 0.00436

1. Calculate K for the following equilibrium when [SO3] = 0.0160M, [SO2] = 0.00560M, and [O2] = 0.00210M. 2SO3(g) ↔ 2SO2(g) + O2(g)

K = [SO2]2 [O2] K = (0.005602 ∙ 0.00210) K = (6.59x10-8) = 1.77x10-5

[SO3]2 ( 0.01602 ) (0.00372)

K<1 concentration of products is larger than reactants…forward reaction favored

1. How would increasing the volume of the reaction vessel affect these equilibria?
   1. NH4Cl(s) ↔ NH3(g) + HCl(g) 0 gases vs. 2 gases…shift right
   2. N2(g) + O2(g) ↔ 2NO(g)2 gases vs. 2 gases…no shift since same amount of gas
2. 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 add away…shift left
   2. removing CO (g) from the system take away towards…shift right
   3. adding H2(g) to the system add away…right
   4. adding something to the system to absorb CO2(g) take away towards…shift left
3. How would simultaneously decreasing the temperature of the system affect these equilibria?
   1. Heat + CaCO3(s) ↔ CaO(s) + CO2(g) take away towards…shift left
   2. 4NH3(g) + 5O2(g) ↔ 4NO(g) + 6H2O(g) + heat take away towards…shift right

**Unit 11 Acids and Bases**

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-

A B CA CB B A CA CB

1. Which is the most concentrated acid:1.0x10-5MHCl or 1.0x10-2M HF largest molarity value
2. Which of the following is a strong acid: 1.0x10-5M HCl one of the 6 strong acids or 1.0x10-2M HF
3. What is the difference between a strong acid or base and a weak acid or base?

Strong acid/base will dissociate or ionize 100%...weak acid/base will dissociate or ionize partially

1. What color would phenolphthalein turn if the solution had a pH of 8.5?pH is basic so pink
2. What color would the litmus paper turn if the solution had a pH of 3.4?pH is acidic so pink
3. Calculate the pH at 298K of a solution having the hydrogen ion concentration of 1.0x10-11M.

[H+] = 1.0x10-11M the exponent is 11 which means pH = 11.0

1. Calculate the pOH and the pH at 298K of a solution having [OH-] = 1.3x10-5M.

[OH-] = 1.3x10-5M the exponent is 5 which means pOH = 5.0 and pOH + pH = 14 so pH = 9.0

1. Label each of the following Acid, Base, Both, or Neutral.
   1. Can turn litmus paper blue Base
   2. Reacts with certain metals Acid
   3. Contains more hydronium ions than hydroxide ions. Acid
   4. Has a pH of 7 Neutral
   5. Feels slippery. Base
   6. Keeps phenolphthalein clear Acid
   7. Tastes bitter. Base
   8. Tastes sour. Base
   9. Is an electrolyte. Both
   10. Can turn litmus paper pink Acid
   11. Occurs when a strong acid reacts with a strong base Neutral
   12. Turns phenolphthalein pink. Base
   13. Contains more hydroxide ions than hydronium ions. Base
   14. Has a pH of 2 Acid
2. 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?

(0.3020M ∙ 33.21mL) = x(20.00mL) 10.03M mL = x(20.00mL) 10.03M mL = x

20.00mL x = 0.502M

1. Write the net ionic equation for copper (II) chloride reacting with ammonium phosphate. List any spectator ions.

3CuCl2(aq) + 2(NH4)3PO4(aq) 🡪 Cu3(PO4)2(s) + 6NH4Cl(aq)

3Cu+2 + ~~6Cl~~~~-~~ + ~~6(NH~~~~4~~~~)~~~~+~~ + 2(PO4)-3🡪 Cu3(PO4)2(s) + ~~6(NH~~~~4~~~~)~~~~+~~ + ~~6Cl~~~~-~~

3Cu+2 + 2(PO4)-3🡪Cu3(PO4)2(s) spectators: (NH4)+ and Cl-

1. Write the net ionic equation for the reaction between barium hydroxide and hydrobromic acid.

List any spectator ions.

Ba(OH)2(aq) + 2HBr(aq) 🡪 BaBr2(aq) + 2H2O(l)

~~Ba~~+2 + 2(OH)- + 2H+ + ~~2Br~~-🡪~~Ba~~+2 + ~~2Br~~- + 2H2O(l)

~~2~~(OH)- + ~~2~~H+🡪~~2~~H2O(l)…(OH)- + H+🡪 H2O spectator ions: Ba+2 and Br-

**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?

Molecules must collide…have enough energy (activation energy)…hit at correct spot

1. What is happening to the bonds when reactions are taking place?

Bonds must be broken during a reaction and then put back together to form a product

1. What happens to the reaction rate when temperature is increased? Explain.

Reaction rate increases because the molecules have more kinetic energy so more collisions have the correct amount of energy (activation energy)

1. What happens to the reaction rate when concentration is increased? Explain.

Reaction rate increases because there is more stuff so there are more collisions

1. What happens to the reaction rate when pressure is increased? Explain.

Reaction rate increases because the molecules are forced closer together so there are more collision

1. What happens to the reaction rate when surface area is increased? Explain.

Reaction rate increases because there are more areas for collisions to occur

1. What happens to the reaction rate when a catalyst is added? Explain.

Reaction rate increases because the catalyst lowers the activation energy so there are more collision with the correct amount of energy (activation energy)

1. What is the sign for ΔH when the reaction is exothermic? Negative Endothermic? Positive
2. 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)

∆T = 79oC – 37oC = 42oC

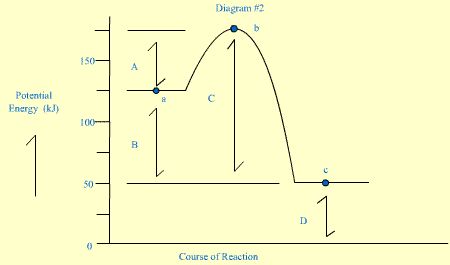
3584J = x (94.8g ∙ 42oC) 3584J = x(3981.6g oC) 3584 J = 0.90 J/goC

3981.6g oC Al (specific heat match)

1. What mass of liquid water at room temperature can be raised to its boiling point with addition of 24kJ of heat energy? Phase change…boiling…so use heat of vaporization

24kJ = 24000J 24000J = (2260J/g)x 24000J = x x = 10.6g

2260J/g

1. How much heat is absorbed when 5.980 grams of water is melted? Phase change…melted…use heat of fusion x = 334J/g ∙ 5.98g x = 1997J
2. Label the potential energy diagram below.

Is this an endothermic or exothermic reaction?

A = activation energy of forward = 175 - 125 = 50kJ

B = heat of reaction, ∆H = 125 – 50 = 75kJ

C = activation energy of reverse = 50 – 175 = -125kJ

D = potential energy of products = 50kJ

Reaction is exothermic because products have lower

Energy than reactions