Notes

Monday, 20 August, 2018

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Significant Figures

Monday, 20 August, 2018

- Numbers 1–9 are always significant.
- Zeros
  - If the 0 is between two non-zero figures, it is significant.
    - *105 = [3]*
    - *10,008 = [5]*
  - If the 0 is at the end of a number with no decimal, it is not significant.
    - *100 = [1]*
    - *28,000 = [2]*
  - If a decimal point is present in the value, the first non-zero and all subsequent digits are significant.
    - *10.00 = [4]*
    - *187.607 = [6]*
    - *200. = [3]*
- Calculations
  - Multiplication and Division
    - Answer will have the same number of significant figures as the value with the least significant figures.
      - $32.6059 \times 0.0031 \times 1.615 = 0.1632414384 \approx 0.16$
    - Round at the end.
  - Addition and Subtraction
    - Answer will have the same number of digits to the right of the decimal point as value with least digits to the right of the decimal point in operation.
      - $59.2356 + 1.053 + 12.1 = 72.3886 \approx 72.4$
    - Round at the end.
  - When combined, round as you go.
    - $\frac{15.617 + 8.02 + 3.153}{2.1} = \frac{26.79}{2.1} \approx \frac{26.79}{2.1}$
      $\frac{26.79}{2.1} = 12.75714286 \approx 13$

Lab Equipment

Wednesday, 22 August, 2018

- Beakers
  - Used to measure estimated quantities of substances, usually liquids
  - Heat resistant
  - Can be used to mix or transport
- Graduated Cylinders
  - Used to measure more exact quantities of liquids
  - Liquids will become concave.
    - Read from eye level at the bottom of the meniscus.
  - No reactions or mixtures
- Flasks
  - Erlenmeyer Flask
    - Use for measuring estimated amounts of liquid and for mixing substances together
    - Ideal for funnels or non-rigorous mixing
    - Short neck
  - Volumetric Flask
    - Used to measure exact amounts of liquids and for mixing substances together
    - Designed for specific measurements
    - For solutions
    - Long narrow neck
  - Round (Florence) Flask
    - Used for heating liquids, usually in a separating process, such as distillation
    - Rarely has a flat bottom
    - Tend to have a frosted glass top
- Burettes
  - Used to titrate various liquids
  - Very controlled distribution of liquids
- Pipettes
  - Used to measure precise amounts of liquids and to carefully transport liquids
  - Two kinds:
    1. Graduated Pipettes
       - Glass
       - Measure
    2. Transfer Pipettes
       - Disposable
       - Plastic
       - Clear
- Funnels
  - Used to separate solids from liquids
  - Can help prevent spills
- Crucibles
  - Used for heating substances
  - Can be used with or without a lid
- Mortars and Pestles
  - Used for grinding down substances into fine particles

- Evaporating Dishes
  - Used for heating up various substances to the point of evaporation
- Watch Glasses
  - Use for crystallization or for beaker covers
  - Slightly concave
  - Can hold samples
- Wire Meshes
  - Used for holding beakers over a flame
  - Makeshift trivet
- Clay Triangle
  - Used to hold a crucible over a flame for heating substances
  - Crucible only rests on top
- Test Tubes
  - Used for holding liquids or specimens, mixing substances together, or for carrying out small scale reactions
  - Can be heat resistant
- Test Tube Racks
  - Used to hold test tubes
- Bunsen Burners
  - Used to heat substances
- Scoopulas
  - Used to transport solids from one container to another
- Stirring Rods
  - Used for stir mixtures
  - Often times glass
- Tweezers
  - Used to pick solids out of other substances to avoid touching or contaminating the substance
- Droppers
  - Used to measure out droplets of liquid
- Balances
  - Used to measure the mass of substances
- Rings
  - Used to support wire mesh or clay triangles which hold other equipment during heating or other processes
- Ring Stands
  - Hold clamps and rings
- Crucible Tongs
  - Used to transport hot crucibles
- Test Tube Tongs
  - Used to transport hot or cold test tubes
- Beaker Tongs
  - Used to transport hot or cold beakers
- Thermometers
  - Used to measure the temperatures of substances
- Well Plates
  - Used to carry out small scale reactions
- Test Tube Clamps or Single Burette Clamps
  - Used to secure test tubes to a stand
  - For heating test tubes

- Burette Clamps
  - Used to secure burettes to a stand
- Googles/Safety Glasses
  - Used to protect eyes from harmful substances
  - Should be worn over prescription glasses
  - Contact lenses should not be worn in the lab

Matter

Thursday, 30 August, 2018

- **Matter**- Anything that has mass and takes up space
- Weight versus Mass
  - Mass is how much matter something has.
  - Weight is the mass multiplied by the acceleration due to gravity.
- States of Matter
  - Changes due to temperature and pressure
  - **Solid**- Definite shape; definite volume
  - **Liquid**- Indefinite shape; definite volume
  - **Gas**- Indefinite shape; indefinite volume
- Changes in Matter
  - Pure Substances
    - **Element**- Something that cannot be broken down further and retain its properties
    - **Compounds**- Two or more elements that are bonded together
      - *i.e.* H_2O; $NaCl$; $C_6H_{12}O_6$
    - Can only be separated chemically
  - Mixtures
    - **Homogenous**- Uniform in composition
      - Solutions
      - Unable to distinguish part
      - *i.e. salt water; air; alloys*
    - **Heterogenous**- Not uniform in composition
      - Able to distinguish parts
      - *i.e. oil and water; chocolate chip cookie dough*
- Separations
  - **Distillations**- Separation due to boiling point differences
  - **Chromatography**- Separation due to differences in size or polarity
  - **Crystallization**- Separation of crystallizing solids out of a solution
- Properties
  - **Physical**- Properties that are observable without altering the identity of the matter
    - *i.e. size, shape, color, boiling point, etc.*
    - **Intensive**- Does not depend on the amount of matter
    - **Extensive**- Does depend on the amount of matter
  - **Chemical**- Properties are not observable without altering the identity of the matter
    - *i.e. flammability, tendency to rust, etc.*
- Changes
  - **Physical**- Alter physical properties
    - *i.e. cutting, changing of state, etc.*
  - **Chemical**- Alter chemical properties
    - *i.e. burning, cooking, etc.*
    - Indicators of Chemical Change
      - Energy Change
        - **Endothermic**- Absorbs energy
        - **Exothermic**- Releases energy
      - Color change
      - Gas formation
      - Formation of a precipitate

Density

Thursday, 30 August, 2018

- $D = \frac{m}{v}$; $D =$ Density; $m =$ mass; $v =$ volume
  - Mass, by itself, and volume, by itself, are both extensive properties, but density is intensive.
  - Density is fixed.
  - $1\text{cm}^3 = 1\text{mL}$
  - $m = Dv$
  - $v = \frac{m}{D}$
- *Practice*
  a. A liquid inside an aquarium has a mass of 2.52kg and the dimensions of the aquarium are 11.25cm × 5.2 cm × 87.92 cm. Calculate the density.
    - $V = Bh = (\ell w)h$
      $= [(11.25\text{cm})(5.2\text{cm})](87.92\text{cm})$
      $= 5143.32\text{cm}^3$
      $D = \frac{m}{v}$
      $= \frac{(2.52\text{kg})}{(5143.32\text{cm}^3)}$
      $= 0.000489955904\,\frac{\text{kg}}{\text{cm}^3}$
      Answer→ 4.9×10^{-4} kilograms per cubic centimeter
  b. A gold bar with a mass of 27.2kg has a density of 19.3g/cm^3. Calculate the volume.
    - $v = \frac{m}{D}$
      $v = \frac{(27.2\text{kg}) = (27200\text{g})}{\left(19.3\,\frac{\text{g}}{\text{cm}^3}\right)}$
      $v = 1409.326425\text{cm}^3$
      Answer→ 1.41×10^3 cubic centimeters

Percent Error

Thursday, 30 August, 2018

- $\%\text{ error} = \left(\frac{|v_e - v_a|}{v_a}\right)100$; $v_e = \text{experimental value}$; $v_a = \text{accepted value}$
- *A soda can states that there are 100g of sugar in the drink. You find there are 98g present. Calculate the percent error.*
  - $\%\ error = \left(\frac{|v_e - v_a|}{v_a}\right)100$
    $= \left[\frac{|(98g) - (100g)|}{(100g)}\right]100$
    $= (0.02)100$
    $= 2\%$

    Answer→ 2.0% error

Atoms: The Building Blocks of Matter

Friday, 14 September, 2018

- The Atom: From Philosophical Idea to Scientific Theory
  - Democritus
    - 'Atom' was based off the Greek word meaning 'invisible.'
  - Foundations of Atomic Theory
    - Laws
      - **Law of Conservation of Mass**- Mass is neither created nor destroyed during ordinary chemical reactions
      - **Law of Definite Properties**- Chemical compounds contain the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound
      - **Law of Multiple Proportions**- If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers
  - Dalton's Atomic Theory
    - John Dalton
    1. All matter is made of atoms.
    2. Atoms of different elements differ in size, mass, and other properties. Atoms of the same element do not differ.
    3. Atoms cannot be subdivided, created, nor destroyed.
    4. Atoms of different elements combine in simple, whole-number ratios to form chemical compounds.
    5. In chemical reactions, atoms are combined, separated, or rearranged.
  - Modern Atomic Theory
    - Atoms are divisible.
    - Atoms of the same element may differ in mass.
      - Isotopes
- The Structure of the Atom
  - **Atom**- The smallest particle of an element that retains the chemical properties of that element
    - 2 Regions
      1) Nucleus
         - Protons
         - Neutrons
      2) Negatively Charged Particles
         - Electrons
  - Discovery of the Electron
    - Cathode Rays and Electrons
      1) Cathode rays were deflected by a magnetic field in the same manner as a wire carrying electric current, which was known to have a negative change.
      2) The rays were deflected away from a negatively charged object.
      - Joseph John Thompson concluded that they were composed of identical negatively charged particles.
        - Named electrons
    - Charge and Mass of the Electron
      1) Because atoms are electronically neutral, they must contain a positive

charge to balance the negative electrons.
2) Because electrons have so much less mass than atoms, atoms must contain other particles that account for most of their mass.

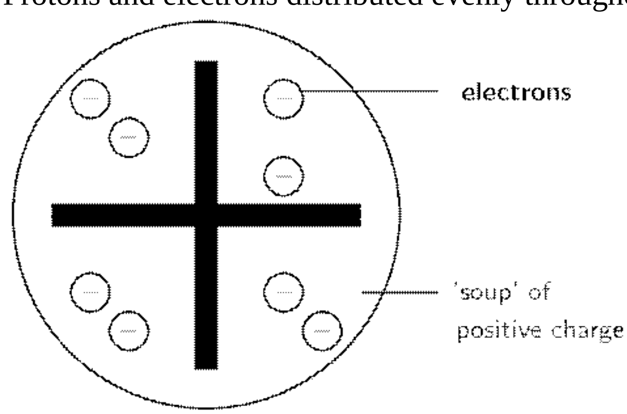
- Discovery of the Atomic Nucleus
  - Rutherford had discovered that the volume of a nucleus was very small compared with the total volume of the atom.
- Composition of the Atomic Nucleus
  - Made of protons and neutrons
  - The number of protons identifies the atom.
  - Forces in the Nucleus
    - As many as 83 protons can exist close together to form a stable nucleus.
    - **Nuclear Forces**- Short-range proton-neutron, proton-proton, and neutron-neutron forces that hold the nucleus particle together
- The Size of Atoms
  - The radius of an atom is the distance from the center of the nucleus to the other portion of the electron cloud.
    - Measure in picometers
      - 1 pm = 10^{-12} m = 10^{-10} cm

- Counting Atoms
  - Atomic Number
    - **Atomic Number (Z)**- The number of protons of each atom of that element
  - Isotopes
    - Hydrogen
      - One proton
      - Isotopes
        - Protium
          - 0 neutrons
          - Most common
        - Deuterium
          - 1 neutron
        - Tritium
          - 2 neutrons
    - **Isotope**- Atoms of the same element that have different masses
    - Tin has 10 stable isotopes.
      - Most of any known element
  - Mass Number
    - **Mass Number**- Total number of protons and neutrons
  - Designating Isotopes
    - Hyphen Notation
      - *Hydrogen-3*
      - *Carbon-14*
    - Nuclear Symbol
      - $^{235}_{92}\mathrm{U}$
      - $^{14}_{6}\mathrm{C}$
      - Superscript indicated mass number.
      - Subscript indicated atomic number.
    - **Nuclide**- A general term for a specific isotope of an element
  - Relative Atomic Mass
    - Grams are too large.
    - Use $^{1}/_{12}$ of the mass of a carbon-12.
      - 1 amu

        - Isotopes may be naturally occuring or may be made in a laboratory.
        - Protons and neutrons are close, but not equal to, 1 amu.
    - Average Atomic Masses of Elements
        - **Average Atomic Mass**- The weighted average of the atomic masses of the naturally occuring isotopes of an element
        - Calculating Average Atomic Mass
            - Average atomic mass of an element depends on mass and relative abundance.
    - Relating Mass to Numbers of Atoms
        - The Mole
            - **Mole (mol)**- The amount of a substance that contains as many particles as there are atoms in exactly 12g of carbon-12
            - Counting unit
        - Avogadro's Number
            - $6.02214179 * 10^{23}$
            - The number of particles in exactly one mole of a pure substance.
        - Molar Mass
            - A mole of carbon-12 atoms has a mass of 12g amd a carbon-12 atom has an atomic mass of 12 amu.
            - **Molar Mass**- The mass of one mole of a pure substance
                - Grams per mole (g/mol)
        - Gram/Mole Conversions
            - *Example*
                - $2.00 \text{ mol He}\left(\frac{4.00\text{g He}}{1 \text{ mol He}}\right) = 8.00\text{g He}$

Atomic Structure

Tuesday, 18 September, 2018

- Early Models of the Atom
  - Democritus
    - *Atomos*
      - Unable to be divided
  - John Dalton
    - Developed an atomic theory
  - Lavoisier
    - Father of Modern Chemistry
  - Dalton's Atomic Theory
    - All elements are composed of indivisible atoms.
      - Atoms break apart into protons, neutrons, and electrons.
    - Atoms of the same element are identical and one different than those of other elements.
      - Isotopes
    - Atoms of different elements can physically mix or chemically combine in simple, whole-number ratios.
    - Chemical reactions occur when atoms are separated, joined, or rearranged. Elements are never changed by chemical changes, only nuclear change.
- Structure of the Atom
  - Subatomic Particles
    - Protons
    - Neutrons
    - Electrons
    - Neutrinos, bosons, quarks, etc.
  - Electrons
    - Negatively charged
    - Discovered by J.J. Thompson using a cathode ray rube
      - Determined charge to mass ratio
      - Robert Millikan determined the mass of an electron.
  - Plum Pudding Model
    - J.J. Thompson
    - Protons and electrons distributed evenly throughout.
    - 

    - Rutherford proved him wrong.
      - Gold Foil Experiment
        - Positive alpha particles should have passed through gold foil.

            - Only some passed through, but some particles were deflected.
        - Must have a positively charged, dense center
            - Nucleus
    -

Particle	Symbol	Relative Electrical Charge	Relative Mass	Actual Mass (g)
Electron	e^-	1-	1/1840 (0)	$9.11 * 10^{-28}$
Proton	p^+	1+	1	$1.67 * 10^{-24}$
Neutron	n^0	0	1	$1.67 * 10^{-24}$

- Distinguishing Between Atoms
    - **__Atomic Number (Z)__**- Number of protons; all defined by their number of protons
    - **__Mass Number__**- The total number of protons and neutrons; relative mass of an atom
    - Atoms tend to be electronically neutral.
        - Same number of protons and electrons
    - **__Isotopes__**- Atoms have the same number of protons, but different numbers of neutrons
        - Chemically alike
        - Differ in physical properties
        - *i.e. carbon-12 or carbon-14*
        - *i.e.* $^{12}_{6}C$ *or* $^{14}_{6}C$ *or* $^{81}_{35}Br^{-}$
    - **__Atomic Mass Unit (amu)__**- $^1/_{12}$ the mass of carbon-12
        - Protons and neutrons are each 1 amu.
        - Average atomic masses are weighted averages of the isotopes.
            - *i.e. Chlorine-35 (75.77% abundance) and chlorine-37 (24.2% abundance) are the two isotopes of chlorine. The periodic table says 35.446 amu.*
            - Skews for the more prevalent isotope.
- **__Ion__**- Atoms that have a charge
    - **__Cation__**- Positive ion; more protons than electrons
    - **__Anion__**- Negative ion; more electrons than protons

Nuclear Chemistry

Monday, 24 September, 2018

- Radioactivity
  - **Nuclear Reactions**- Reactions in which the nuclei of unstable
  - History
    - Henri Becquerel noticed that uranium ore exposed photographic filter.
    - Marie and Pierre Curie named the process 'radioactivity.'
    - The particles and rays emitted by a radioactive source called 'radiation.'
  - Disproves Dalton's Theory of Invisible Atoms
  - Stability depends on the proton-neutron ratio and general size.
    - Undergoes radioactive decay to become more stable
    - More protons = More repulsion forces
- Types of Radiation
  - Alpha Radiation
    - Helium nucleus = Alpha particle
      - 2 protons and 2 neutrons
        - No electrons
      - ${}^{4}_{2}\mathrm{H}$ or ${}^{4}_{2}\alpha$
    - Lowest energy radiation
  - Beta Radiation
    - Electrons = Beta particles
      - ${}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$
    - Numbers of neutrons decrease and protons increase.
    - A neutron becomes a proton and an electron.
      - Mass stays the same.
      - The electron is ejected.
  - Gamma Radiation
    - ${}^{0}_{0}\gamma$
    - High energy electromagnetic radiation
    - In conjunction with alpha and beta radiation
    - Most dangerous
      - Interferes with cells
  -

Property	**Alpha**	**Beta**	**Gamma**
Composition	Helium Nucleus Alpha Particle	Electron Beta Particle	Photon Gamma Particle
Symbol	${}^{4}_{2}\mathrm{H}$ or ${}^{4}_{2}\alpha$	${}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$	${}^{0}_{0}\gamma$
Charge	2+	1-	0
Mass	4	1/1837	0
Shielding	Paper, clothing	Metal fork	Lead, concrete

- Nuclear Stability and Decay
  - 264 of the 500 nuclei do not decay with time.
  - Elements of a low atomic number (less than 20), stable nuclei have roughly equal protons and neutrons.
  - In higher atomic numbers, there needs to be more neutrons.
  - Too many neutrons

        - Beta decay
    - Too few neutrons
        - Electron capture
            - Opposite of beta decay
    - Positron
        - A particle with the mass of an electron, but has a positive charge.
        - ${}^{0}_{+1}e$

Half-Life

Wednesday, 26 September, 2018

- Every radioisotope has a rate of decay.
- **Half-Life**- Time required for one-half of the nuclei of a radioactive sample to decay products
- Transmutation Products
    - **Transmutation**- Conversion of an atom of one element to an atom of another element
    - Radioactive Decay
    - Bombarding the nucleus with high energy particles
    - All elements with an atomic number above 92 are transmutation elements.
        - Do not occur in nature
            - Synthesized in nuclear reactors
        - Radioactive
- Nuclear Fission
    - A nucleus is bombarded with particles, typically a neutron, and the nucleus splits into smaller fragments.
    - Generate additional neutrons
        - Chain reactions
    - Generate a lot of energy
        - Radioactive waste
- Nuclear Fusion
    - Nuclei combine to produce a nucleus of greater mass.
    - Requires high temperatures
        - Matter exists as plasma at that heat.
        - 6x the heat of the core of the sun
- *Practice*
    a. The half-life of cesium-137 is 30.2 years. If the initial mass of the sample is 1.00 kg, how much will remain after 151 years?
        - $\frac{\text{Time Elapsed}}{\text{Half-Life}} = n = \frac{151\text{ years}}{30.2\text{ years}} = 5\text{ half-lives}$
        $(\text{Initial Mass})\left(\frac{1}{2}\right)^{n} = (1.00\text{ kg})\left(\frac{1}{2}\right)^{(5)} = 0.03125\text{ kg}$
    Answer→ $3.13 * 10^{-2}$ kilograms
    b. A 64 kg sample of germanium-66 is left undisturbed for 12.5 hours. At the end of that period, only 2.0 g remain. What is the half-life of this material?
        - $\frac{\text{Initial Mass}}{\text{Final Mass}} = 2^{n} = \frac{64\text{ kg}}{2.0\text{ kg}} = 5$
        $\frac{(12.5\text{ hours})}{x} = (5)$
        $x = 2.5\text{ hours}$
    Answer→ 2.5 hours

Energy, Frequency, and Wavelength

Tuesday, 2 October, 2018

- Light and Atomic Spectra
  - Isaac Newton proposed light was made of particles.
  - Scientists accepted light as a wave.
    - Light consists of electromagnetic waves.
      - Radio Waves
      - Microwaves
      - Infrared Rays
      - Visible Light
      - Ultraviolent Rays
      - X-Rays
      - Gamma Rays
      - Cosmic Rays
  - 4 Properties for Measuring Waves
    i. **Amplitude**- Wave's height from the origin to the crest
    ii. **Wavelength (λ)**- Distance between crests
    iii. **Frequency (ν)**- The number of wave cycles to pass a given point per unit of time
      - Cycles per second
      - Hertz (Hz) = s^{-1}
    iv. **Speed (c)**- The rate at which the wave travels
      - $c = \lambda \nu$
        - $c = 3.00 * 10^8 \frac{\mathrm{m}}{\mathrm{s}}$
  - Energy element emits light when it is excited by the passage of an electric discharge through its gas or vapor.
  - Passing light through a prism results in the atomic emission spectrum of the element.
- Quantum Concept
  - Classical physics does not explain the line spectra of the elements.
  - Max Planck
    - Why iron changes color when heated
    - $E = h\nu$
      - $h = 6.626 * 10^{-54}\mathrm{Js}$
        - Planck's Constant
  - Energy absorbed or emitted is a quantum (quanta).
    - Quanta are very small and have fixed increments.

Photoelectric Effect

Thursday, 4 October, 2018

- When light is shined on the surface of a metal, electron are ejected from the metal atoms.
    - Light of a certain frequency
    - Acts as a particle
        - Photon
        - Dual nature of light

Electron Configuration

Thursday, 4 October, 2018

- Energy Levels
    - 1 – 7
    - Distance from the nucleus
    - Broken into sublevels
        - *s*
        - *p*
        - *d*
        - *f*
        - Broke down into orbitals
- **Aufbau Principle**- Fill lowest available energy levels first
- **Pauli Exclusion Principles**- There are a maximum of 2 electrons per orbital and must have opposite spins
- **Hund's Rule**- Electron spread over the orbitals in a sublevel before making pairs

Periodic Table

Tuesday, 23 October, 2018

- Demetri Mendeleev
    - Arranged elements by similarities in properties and mass
    - Left blank spaces
- Henry Moseley arranged elements by atomic number.
- Rows are called periods.
- Columns are called groups or families.
- **Periodic Law**- When elements are arranged by atomic number; there is a periodic repetition of physical and chemical properties
- Groups
    - Group 1: Alkali Metals
    - Group 2: Alkaline Earth Metals
    - Group 3–12: Transition and Inner Transition Metals
    - Group 17: Halogens
    - Group 18: Noble Gases/Inert Gases
- Most elements are solid at 25 °C (room temperature).
    - Mercury and bromine are liquids.
    - The noble gases, hydrogen, nitrogen, oxygen, fluorine, and chlorine are gaseous.
- All elements above 92, technetium, and promethium are man-made.
- Classification of Elements
    - Chemical properties are due to electron configuration.
    - **Representative Elements**- Outermost *s* and *p* sublevels are only partially filled; elements within the *s*-block or the *p*-block
- Periodic Trends
    - **Atomic Radius**- Half the distance between two nuclei of a diatomic molecule
        - Increases down a group
        - Decreases across a period
        - **Ionic Radius**- If ions are not being compared to their original atoms and are not isotopes, the trend is the same as the atomic radius
            - Cations are generally smaller.
            - Anions are generally larger.
            - **Isoelectric Ions**- When ions have the same number of electrons; the ions with the more protons will be smaller
            - Due to Coulomb's Law
                - $F = k\left(\frac{Q_1 Q_2}{d^2}\right)$
                    - $F =$ Force
                    - $k =$ Planck's Constant
                    - $Q =$ Charge of the particle
                    - $d =$ distance between the centers of the particles
    - **Ionization Energy**- Energy required to remove an electron
        - Decreases down a group
        - Increases across a period
    - **Electronegativity**- The tendency for atoms of an element to attract electrons in a chemical bond
        - Decreases down a group

    - Increases across a period
    - Same trend as ionization energy
    - Noble gases have no electronegativity.
  - Metallic-ness increases downward and leftwards.

Naming Ions

Thursday, 8 November, 2018

- **Cation**- Positively charged ion
  - *Element name* + "ion"
  - *Examples*
    - Na^{+}: Sodium ion
    - Ca^{2+} : Calcium ion
- **Anion**- Negatively charged ion
  - Change element ending to "ide."
  - *Examples*
    - Cl^{2-}: Chloride (ion)
    - O^{2-}: Oxide (ion)
- **Ionic Bond**- Transfer of electrons between a metal and a nonmetal or polyatomic ion
  - Nomenclature
    - Name the cation then the ion and drop the "ion."
    - *Examples*
      - NaCl: Sodium ion + Chloride ion = Sodium chloride
      - K_2O: Potassium ion + Oxide ion = Potassium oxide
      - Al_2O_3: Aluminum ion + Oxide ion = Aluminum oxide
      - Polyatomic Ions
        - $CaSO_4$: Calcium ion + Sulfate = calcium sulfate
  - Name to Formula
    - Charges must balance
      - *Examples*
        - Lithium nitride
          - Li^{+} and N^{3-}
          - Li_3N
        - Aluminum sulfide
          - Al^{3+} and S^{2-}
          - Al_2S_3
        - Sodium sulfate
          - Na^{+} and SO_4^{2-}
          - Na_2SO_4
        - Aluminum sulfate
          - Al^{3+} and SO_4^{2-}
          - $Al_2(SO_4)_3$
    - If the ratio can be reduced, reduce it.
- Multivalent Electrons
  - *CrI_2 versus CrI_3*
    - *Both are correct.*
  - Use Roman numerals to indicate the charge
    - *Chromium(II) iodide versus Chromium(III) iodide*
  - Latin Affixes

Element	Stem
Iron (Fe)	Ferr-
Copper (Cu)	Cupr-

Gold (Au)	Aur-
Lead (Pb)	Plumb-
Tin (Sn)	Stann-

- Suffix
  - -ic
    - Higher charge
  - -ous
    - Lower charge
- *Examples*
  - Plumbic carbonate
    - Lead(IV) carbonate
    - $Pb_2(CO_3)_4 = Pb(CO_3)_2$
  - Tin(IV) oxide
    - Stannic oxide
    - $Sn_2O_4 = SnO_2$
  - Zinc nitride
    - Zinc is not multivalent.
    - Zn_3N_2

Covalent Bonding

Wednesday, 14 November, 2018

- Takes place between two nonmetals and involves a sharing of electrons to obtain an octet
- Electrons that do not take place in the bond are called unshared pairs.
- Sometimes atoms will form multiple bonds to achieve stable configurations.
    - Single, double, or triple bonds
    - Single bonds are the longest, but triple is the strongest.
- Diatomic elements and polyatomic ions are covalently bonded.
    - BrINClHOF
        - Bromine
        - Iodine
        - Nitrogen
        - Chlorine
        - Hydrogen
        - Oxygen
        - Fluorine
- Nomenclature
    - **Molecules**- Covalently bonded compounds
    -

Number	**Prefix**
1	Mono-
2	Di-
3	Tri-
4	Tetra-
5	Penta-
6	Hexa-
7	Hepta-
8	Octa-
9	Nona-
10	Deca-

    - Both elements in the compound need a prefix unless the first one is mono.
    - *Examples*
        - H_2O: Dihydrogen monoxide
        - CO_2: Carbon dioxide

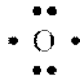
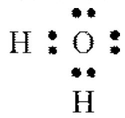
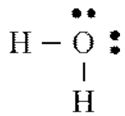
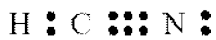
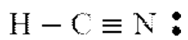
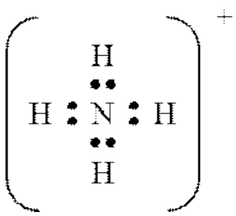
Naming Acids and Hydrates

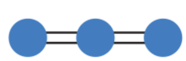
Wednesday, 14 November, 2018

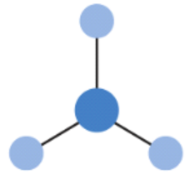
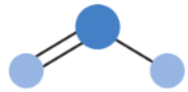
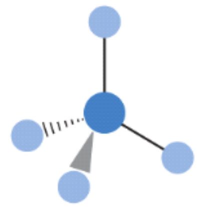
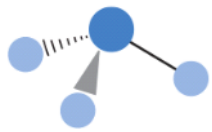
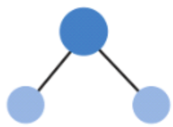
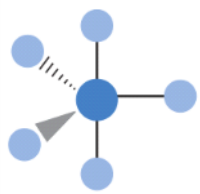
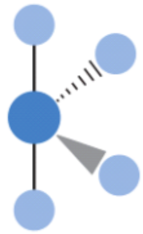
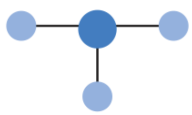
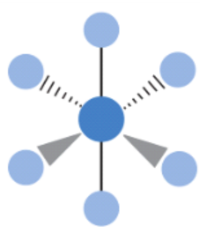
- Acids
    - Acids contain hydrogen ions.
    - If it is bounded to a monatomic ion, the prefix "hydro-" is added and the ending becomes "ic."
    - Polyatomic Ions
Polyatomic Ion	Acid
Hypo _____ ite	Hypo _____ous acid
_____ ite	_____ ous acid
_____ ate	_____ ic acid
Per _____ ate	Peri _____ ic acid
    - *Examples*
        - HCl: Hydrochloric acid
        - HNO_3: Nitric acid
        - HF: Hydrofluoric acid
        - H_2SO_4: Sulfuric acid
- Hydrates
    - Include H_2O in their structure
    - Name the base portion first, and then name the hydrate portion.
    - *i.e. $CuSO_4 \bullet 5H_2O$ is cupric sulfate pentahydrate or copper(II) pentahydrate.*

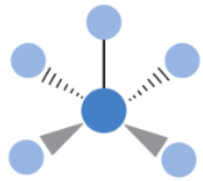
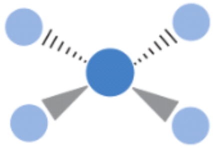
Lewis Dot Diagrams and Molecular Shape

Friday, 16 November, 2018

- Lewis Dot Diagrams
    - *Examples*
        - Oxygen
            - 6 valence electrons
            - 
        - Water
            - H_2O
            - 8 valence electrons
            - 
            - 
        - HCN
            - 10 valence electrons
            - 
            - 
        - Ammonium
            - NH_4^+
            - 8 valence electrons
                - Minus 1 due to charge
            - 
- Molecular Shape
    - Dependent on position of the electrons
    - VSEPR Theory
        - Valence Shell Electron Pair Repulsion
        - Electrons in compounds will arrange themselves as far apart as possible.

Bonding Groups	**Lone Pairs**	**Molecular Geometry**	**Shape**	**Notes**	**Examples**
2	0	Linear		• Between 2 or 3 atoms • Includes a multiple bond • Bond angle = 180°	HBr HCN CO_2 CS_2

3	0	Trigonal Planar		• Bond angle = 120°	BCl_3 $AlCl_3$
2	1	Bent		• Bond angle = 105°	H_2O SO_2
4	0	Tetrahedral		• Bond angle = 109.5°	CH_4 CCl_4
3	1	Trigonal Pyramidal		• Bond angle = 107°	NH_3 PH_3
2	2	Bent		• Bond angle = 105°	H_2S
5	0	Trigonal Bipyramidal		• Bond angle (equatorial) = 120° • Bond angle (axial) = 90°	PCl_5
4	1	Seesaw		• Bond angle (equatorial) = <120° • Bond angle (axial) = <90°	SF_4
3	2	T-Shaped		• Bond angle = <90°	ClF_3
2	3	Linear		• Bond angle = 180°	XeF_2
6	0	Octahedral		• Bond angle = 90°	SF_6

5	1	Square pyramidal		• Bond angle = <90°	BrF_5
4	2	Square planar		Bond angle = 90°	XeF_4

Properties of Molecules

Tuesday, 20 November, 2018

- Bonds can either be polar or nonpolar.
  - **Polar Covalent Bond**- Electrons are shared unequally
  - **Non-Polar Covalent Bond**- Electrons are shared equally
  - Determined by electronegativity differences
    - 0.0 – 0.4 is non-polar covalent.
    - >0.4 – 2.0 is polar covalent.
    - >2.0 is ionic.
- **Dipole**- Polar molecule
  - Depends on:
    - The presence of polar bonds
    - Molecular shape
      - Asymmetrical molecules
- *Examples*
  a. H_2O is polar.
    - Bent shape
      - Asymmetrical
    - Net dipole vector towards the oxygen atom.
    - The oxygen has a partial negative (δ-) charge and the hydrogen have a partial positive (δ+) charge.

  b. CO_2 is not polar.
    - Carbon and oxygen have a polar bond.
    - Linear shape
      - Symmetrical
    - Dipole vector cancel out.
      - Are equal in opposite directions

Metallic Bonding

Tuesday, 27 November, 2018

- Between two metals
- Electrons are denuclearized
  - "Sea of electrons"
- Properties of Ionic Compounds
  - Hard
  - Brittle
  - Have high boiling/melting points
  - Conduct electricity when molten or in solution
- Properties of Molecules
  - Polar or nonpolar
- Metals
  - Shiny
    - Luster
  - Malleable
  - Ductile
  - Flexible due to the sea of electrons
  - Alloys are mixtures of metals.

Moles

Monday, 3 December, 2018

- **Mole**- SI unit of quantity/amount
  - 6.022×10^{23} particles per mole
    - Avogadro's Number
- Representative Particles
  - Element: Atom
  - Covalent Compound: Molecule
  - Ionic Compound: Formula Unit
- Conversion
  - Particles ↔ Mole
    - Avogadro's Number
  - Mass ↔ Mole
    - Molar mass
  - Volume ↔ Mole
    - For gases at STP
      - 0 °C and 1.00 atm
    - 22.4 liters in one mole

Percent Composition, Empirical Formula, and Molecular Formula

Wednesday, 5 December, 2018

- Percent Composition
  - **<u>Percent Composition</u>**- Finding the percent by mass of an element in a compound
  - For each element, divide the molar mass of the element by the total mass of the compound and convert to a percent.
  - *Example*
    i. Determine the percent composition of sodium oxalate.
      - Total mass: 134.0 amu
      - Sodium
        - $\frac{2(22.99)}{134.0} = 0.3431343284 = 34.31\%$
      - Carbon
        - $\frac{2(12.01)}{134.0} = 0.1792537313 = 17.93\%$
      - Oxygen
        - $\frac{4(16.00)}{134.0} = 0.4776119403 = 47.76\%$

    Answer→ $Na_2C_2O_4$: Na – 34.31%; C – 17.93%; O – 47.76%

    ii. Calculate the mass of the given element in the following compound. Bromine in 50.0 g potassium bromide.
      - $50.0 \text{ g KBr}\left(\frac{1 \text{ mol KBr}}{119 \text{ mol KBr}}\right)\left(\frac{1 \text{ mol Br}}{1 \text{ mol KBr}}\right)\left(\frac{79.9 \text{ g Br}}{1 \text{ mol Br}}\right)$
        $= 33.57142857 \text{ g Br}$

    Answer→ 33.6 grams bromine
- Empirical Formula
  - **<u>Empirical Formula</u>**- Shows the lowest whole number ratio for elements in a compound
  - If the chemical formula is already known, the subscripts in the chemical formula can be simplified to the lowest whole number ratio.
    - *i.e.* $C_6H_{12}O_6 \rightarrow CH_2O$
  - If the chemical formula is not known:
    i. If the components are already given in grams, proceed to step ii. If the components are given in percents, assume there is a 100 gram sample (essentially drop the percent symbol and make gram).
    ii. Convert each component to moles.
    iii. Divide all of the moles by the smallest quantity of moles.
    iv. The numbers of step iii become the subscripts of the empirical formula (if they are whole numbers of relatively close/within a few hundredths). If a third or two thirds, multiply all numbers by 3. If a half, multiply all numbers by 2.
  - *Example*
    - Determine the empirical formula for a compound that is 60.9% As and 39.1% S.
      1) 60.9% As → 60.9 g As
         39.1% S → 39.1 g S
      2) 60.9 g As → 0.8129 mol As
         39.1 g S → 1.219 mol S
      3) 0.8129 mol As → 1.000

1.219 mol S → 1.500

4) $AsS_{1.5} \rightarrow As_2S_3$

Answer→ Diarsenic trisulfide (As_2S_3)

- Molecular Formula
  - **Molecular Formula**- Shows the types and quantities of atoms in a compound
  - Need to know the empirical formula and the molecular/molar mass
  - Steps
    - i. Divide the molecular mass by the empirical formula mass (this number should be a whole number or within a few hundredths).
    - ii. Multiply the subscripts in the empirical formula by the number from step i.
  - *Example*
    - Determine the molecular formula for a compound with an empirical formula of C_2H_4S and a molar mass of 179 amu.
      - Empirical formula mass: 60.125 amu

1) $\frac{179}{60.125} = 2.98$

2) $C_6H_{12}S_3$

Answer→ $C_6H_{12}S_3$

2-Step Mole Conversions and Molarity

Friday, 7 December, 2018

- 2-Step Mole Conversions
  - If converting from particles/grams/liters to particles/grams/liters, you need to go through moles first.
  - *Example*
    - Convert 15.9 g NaCl to formula units of NaCl.
      - $15.9 \text{ g NaCl}\left(\frac{1 \text{ mol}}{58.44 \text{ g}}\right)\left(\frac{6.022 \times 10^{23} \text{ fu}}{1 \text{ mol}}\right) = 1.638429158 \times 10^{23} \text{ fu NaCl}$

      Answer→ 1.64×10^{23} formula units of sodium chloride
  - Significant figures still match the number of significant figures in original given quantity.
  - Conversion factos are still set up so that units still cancel out that the values and units match up as follows:
    - mol = 1
    - L = 22.4
    - g = molar mass
    - particles 6.022×10^{23}
- Molarity
  - **<u>Molarity (M)</u>**- A quantitative measure of concentrations of a solution; moles per liter
  - $M = \frac{\text{moles of solute}}{\text{liters of solution}}$
  - *Example*
    - Calculate the molarity when 42 g $MgCl_2$ is in 425 mL of water.
      - $\frac{42 \text{ g } MgCl_2}{425 \text{mL } H_2O}\left(\frac{1 \text{ mol } MgCl_2}{95.21 \text{ g } MgCl_2}\right)\left(\frac{1000 \text{ mL } H_2O}{1 \text{ L } H_2O}\right) = 1.037953255 \frac{\text{mol}}{\text{L}}$

      Answer→ 1.0 molar magnesium chloride

Chemical Reactions

Tuesday, 8 January, 2019

- Chemical reactions are expressed with chemical equations.
- Word Equations
  - Reactants → Products
  - *Iron + Oxygen → Iron(III) Oxide*
- **Skeleton Equations**- Equations that only show the formulas of the reactants and products
  - $Fe + O_2 \rightarrow Fe_2O_3$
- Include states of substances
  - Not critical
  - **Aqueous**- In a solvent
- Balancing Equations
  - Obey the Law of Conservation of Matter
  - $\mathrm{Fe} + \mathrm{O_2} \rightarrow \mathrm{Fe_2O_3}$
1 Fe	2 Fe
2 O	3 O

      $\mathrm{4Fe} + \mathrm{3O_2} \rightarrow \mathrm{2Fe_2O_3}$

    Answer→ $\mathrm{4Fe} + \mathrm{3O_2} \rightarrow \mathrm{2Fe_2O_3}$
  - $\mathrm{Na} + \mathrm{Cl_2} \rightarrow \mathrm{NaCl}$

    Answer→ $\mathrm{2Na} + \mathrm{Cl_2} \rightarrow \mathrm{2NaCl}$
  - $\mathrm{CH_4} + \mathrm{O_2} \rightarrow \mathrm{CO_2} + \mathrm{H_2O}$

    Answer→ $\mathrm{CH_4} + \mathrm{2O_2} \rightarrow \mathrm{CO_2} + \mathrm{2H_2O}$
  - $\mathrm{C_8H_{18}} + \mathrm{O_2} \rightarrow \mathrm{CO_2} + \mathrm{H_2O}$
    - $\mathrm{C_8H_{18}} + \frac{25}{2}\mathrm{O_2} \rightarrow \mathrm{8CO_2} + \mathrm{9H_2O}$

    Answer→ $\mathrm{2C_8H_{18}} + \mathrm{25O_2} \rightarrow \mathrm{16CO_2} + \mathrm{18H_2O}$
  - Aluminum sulfate + lead (II) nitrate → Aluminum nitrate + lead (II) sulfate
    - $\mathrm{Al_2(SO_4)_3} + \mathrm{Pb(NO_3)_2} \rightarrow \mathrm{Al(NO_3)_3} + \mathrm{PbSO_4}$

    Answer→ $\mathrm{Al_2(SO_4)_3} + \mathrm{3Pb(NO_3)_2} \rightarrow \mathrm{2Al(NO_3)_3} + \mathrm{3PbSO_4}$

Predicting Reactions

Thursday, 10 January, 2019

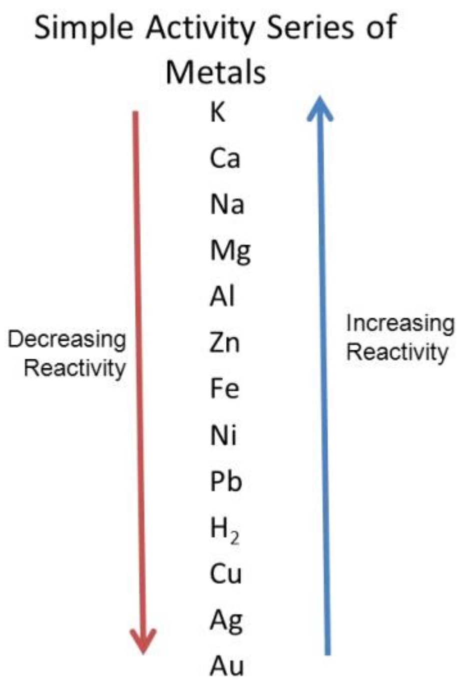
1. Decomposition: $AB \rightarrow A + B$
2. Combustion: $\mathrm{C}_x\mathrm{H}_y + \mathrm{O}_2 \rightarrow \mathrm{CO}_2 + \mathrm{H}_2\mathrm{O}$
3. Double Displacement: $AB + XY \rightarrow AY + XB$
4. Single Displacement: $AX + Y \rightarrow YX + A$
5. Synthesis: $A + B \rightarrow AB$

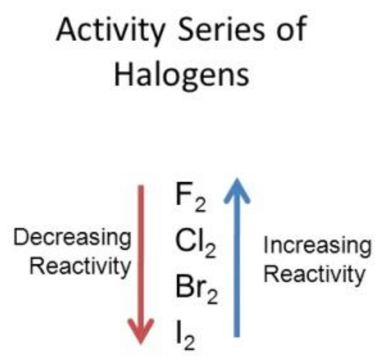
- How to predict the type of reaction from the reactants:
    -

		Yes	*No*
1	Is there one reactant?	Decomposition	2
2	Is one of the reactants an organic compound?	Combustion	3
3	Is one of the reactants an element?	4	Double displacement
4	Is one of the reactants H_2O?	Single displacement	5
5	Is one of the reactants O_2 or CO_2?	Synthesis	Single Displacement

- Decomposition Rules
    - All binary compounds → Elements
        - $H_2O \rightarrow H_2 + O_2$
        - $MgCl_2 \rightarrow Mg + Cl_2$
    - All carbonates → The oxide + CO_2
        - $CaCO_3 \rightarrow CaO + CO_2$
        - $NaCO_3 \rightarrow Na_2O + CO_2$
    - Chlorates → Binary salt + oxygen
        - $KClO_3 \rightarrow KCl + O_2$
        - $Ba(ClO_3)_2 \rightarrow BaCl_2 + O_2$
    - Bases → Oxide of the metal + water
        - $Ca(OH)_2 \rightarrow CaO + H_2O$
        - $NaOH \rightarrow Na_2O + H_2O$
    - Acids → oxide of the nonmetal
        - $HNO_3 \rightarrow N_2O_5 + H_2O$
        - *(acid)* $H_3PO_4 \rightarrow P_2O_5 + H_2O$
- Combustion Rules
    - $\mathrm{C}x\mathrm{H}y + \mathrm{O}_2 \rightarrow \mathrm{H}_2\mathrm{O} + \mathrm{CO}_2$
    - $\mathrm{C}x\mathrm{H}y\mathrm{N}z + \mathrm{O}_2 \rightarrow \mathrm{H}_2\mathrm{O} + \mathrm{CO}_2 + \mathrm{NO}_2$
    - $\mathrm{C}x\mathrm{H}ySz + \mathrm{O}_2 \rightarrow \mathrm{H}_2\mathrm{O} + \mathrm{CO}_2 + \mathrm{SO}_2$
- Double Displacement
    - Anions switch places and cation switch places
- Single Displacement
    - Decide if element normally forms a positive or negative ion.
    - If negative ion, anion from reactant switches with it
    - If positive ion, cation switches with it
- Synthesis
    - Ask yourself what type of decomposition produces these products

- For Single and Double Displacement (continued list):

    - Union of element → Binary compounds
    - Metallic oxide + carbon dioxide → Carbonates
    - Binary salt + oxygen → Chlorates
    - Metallic oxide + water → Base
        - $CaO + H_2O \rightarrow Ca(OH)_2$
        - $Na_2O + H_2O \rightarrow NaOH$
    - Non-metallic oxide + water → Acid
        - $N_2O_5 + H_2O \rightarrow HNO_3$
        - $P_2O_5 + H_2O \rightarrow H_3PO_4$
- For Single and Double Displacement
    - 



☺

    - **Table F**
Solubility Guidelines for Aqueous Solutions

Ions That Form *Soluble* Compounds	Exceptions
Group 1 ions (Li^+, Na^+, etc.)	
ammonium (NH_4^+)	
nitrate (NO_3^-)	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	
hydrogen carbonate (HCO_3^-)	
chlorate (ClO_3^-)	
perchlorate (ClO_4^-)	
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}

Ions That Form *Insoluble* Compounds	Exceptions
carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
chromate (CrO_4^{2-})	when combined with Group 1 ions, Ca^{2+}, Mg^{2+}, or ammonium (NH_4^+)
phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}, or ammonium (NH_4^+)

Oxidation-Reduction Reactions

Monday, 14 January, 2019

- Paired process
- Transfer of electrons
- OIL RIG
- Use oxidation numbers
- Assigning Oxidation Numbers
  - An uncombined element = 0
  - Monatomic ion = Charge on the ion
  - Fluorine = –1
  - Oxygen = –2
    - In fluorine, it's +1 or +2.
    - In peroxides, it's –1.
  - Hydrogen = +1
    - In a metal = –1
  - Group 1 = +1
  - Group 2 = +2
  - Aluminum = +3
  - Sum of all in a compound = 0
  - Sum of all in a polyatomic ion = Charge on the ion
  - Examples
    - i. H_2O
      Answer→ $\overset{+1}{H}_2\overset{-2}{O}$
    - ii. $HCLO_4$
      Answer→ $\overset{+1}{H}\overset{+7}{Cl}\overset{-2}{O}_4$
    - iii. SO_4^{-2}
      Answer→ $\overset{+6}{S}\overset{-2}{O}_4^{-2}$
    - iv. NO_3^-
      Answer→ $\overset{+5}{N}\overset{-2}{O}_3^-$
    - v. O_2
      Answer→ $\overset{0}{O}_2$
    - vi. $AlCl_3$
      Answer→ $\overset{+3}{Al}\overset{-1}{Cl}_3$
    - vii. MnO_4^-
      Answer→ $\overset{+7}{Mn}\overset{-2}{O}_4^-$
    - viii. $AgNO_3$
      Answer→ $\overset{+1}{Ag}\overset{+5}{N}\overset{-2}{O}_3$
- Determining if a redox reaction has occurred
  - $Cu + AgNO_3 \rightarrow Cu(NO_3)_2 + Ag$
    - If oxidation number has increased, then it has been oxidized.
    - Copper is oxidized; silver is reduced.
    - Must have both!
  - $H_2O_2 \rightarrow H_2O + O_2$
    - Oxygen is both oxidized and reduced.
  - $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$
    - Not a redox reaction

Net Ionic Equations

Friday, 18 January, 2019

- Complete Ionic Equation
  - $Cu(s) + 2Ag^{+} + 2NO_3^{-} \rightarrow Cu^{2+} + 2NO_3^{-} + 2Ag$
    - Nitrate is not reacting.
      - Spectator ion
    - Net Ionic Equation
      - $Cu + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag$
- Examples
  - $Cu + NO_3^{-} \rightarrow Cu^{2+} + NO$
    -

Copper is oxidized.	Nitrogen is reduced.
$Cu \rightarrow Cu^{2+}$	$NO_3^{-} \rightarrow NO$
$Cu \rightarrow Cu^{2+}$	$NO_3^{-} \rightarrow NO$
$Cu \rightarrow Cu^{2+}$	$NO_3^{-} + 4H^{+} \rightarrow NO + 2H_2O$
$Cu \rightarrow Cu^{2+} + 2e^{-}$	$NO_3^{-} + 4H^{+} + 3e^{-} \rightarrow NO + 2H_2O$
$3Cu \rightarrow 3Cu^{2+} + 6e^{-}$	$2NO_3^{-} + 8H^{+} + 6e^{-} \rightarrow 2NO + 4H_2O$

Answer$\rightarrow$ $3Cu + 2NO_3^{-} + 8H^{+} \rightarrow 2Cu^{2+} + 2NO + 4H_2O$

  - $Zn + HCl \rightarrow Zn^{2+} + H_2$
    -

Zinc is oxidized.	Hydrogen is reduced.
$Zn \rightarrow Zn^{2+}$	$HCl \rightarrow H_2$
$Zn \rightarrow Zn^{2+}$	$HCl \rightarrow H_2 + Cl^{-}$
$Zn \rightarrow Zn^{2+}$	$HCl + H^{+} \rightarrow H_2 + Cl^{-}$
$Zn \rightarrow Zn^{2+} + 2e^{-}$	$HCl + H^{+} + 2e^{-} \rightarrow H_2 + Cl^{-}$

Answer$\rightarrow$ $Zn + HCl + H^{+} \rightarrow Zn^{2+} + H_2 + Cl^{-}$

- Basic Reactions
  - If it is basic, use OH^{-} to balance out H^{+}.
    - Must add to both sides.
  - $H^{+} + OH^{-} \rightarrow H_2O$

Kinetics

Tuesday, 29 January, 2019

- **Collision Theory**- Molecules must collide in order to react
- Rate Law
  - rate $= k[\mathrm{A}]^x[\mathrm{B}]^y$

Stoichiometry

Friday, 8 February, 2019

- Examples
  a. $2H_2 + O_2 \rightarrow 2H_2O$
     - **<u>Mole Ratio</u>**- Comes from coefficients of a balanced equation
     - If 4.5 moles of O_2 reacts with excess H_2, how many moles of H_2O are produced?
       - $4.5\ \text{mol}\ \text{O}_2 \left(\frac{2\ \text{mol}\ \text{H}_2\text{O}}{\text{mol}\ \text{O}_2}\right) = 9\ \text{mol}\ \text{H}_2\text{O}$

     Answer→ 9.0 moles of water
  b. If 4.5 grams of O_2 reacts with excess H_2, how many moles of H_2O are produced?
     - $4.5\ \text{g}\ \text{O}_2 \left(\frac{\text{mol}\ \text{O}_2}{32\ \text{g}\ \text{O}_2}\right)\left(\frac{2\ \text{mol}\ \text{H}_2\text{O}}{\text{mol}\ \text{O}_2}\right) = 0.28\ \text{mol}\ \text{H}_2\text{O}$

     Answer→ 0.28 moles of water
  c. If 4.5 grams of O_2 reacts with excess H_2, how many grams of H_2O are produced?
     - $4.5\ \text{g}\ \text{O}_2 \left(\frac{\text{mol}\ \text{O}_2}{32\ \text{g}\ \text{O}_2}\right)\left(\frac{2\ \text{mol}\ \text{H}_2\text{O}}{\text{mol}\ \text{O}_2}\right)\left(\frac{18\ \text{g}\ \text{H}_2\text{O}}{\text{mol}\ \text{H}_2\text{O}}\right) = 5.1\ \text{g}\ \text{H}_2\text{O}$

     Answer→ 5.1 grams of water
- Assume excess if there is no amount given.
- Limiting Reactant Stoichiometry
  - **<u>Limiting Reactant</u>**- Reactant used up first in a reaction
  a. Identify the limiting reactant when 1.22 grams O_2 react with 1.05 grams H_2 to produced H_2O?
     - $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
     - $1.22\ \text{g}\ \text{O}_2 \left(\frac{\text{mol}\ \text{O}_2}{32\ \text{g}\ \text{O}_2}\right)\left(\frac{2\ \text{mol}\ \text{H}_2\text{O}}{\text{mol}\ \text{O}_2}\right)\left(\frac{18\ \text{g}\ \text{H}_2\text{O}}{\text{mol}\ \text{H}_2\text{O}}\right) = 1.37\ \text{g}\ \text{H}_2\text{O}$
     - $1.05\ \text{g}\ \text{H}_2 \left(\frac{\text{mol}\ \text{H}_2}{2\ \text{g}\ \text{H}_2}\right)\left(\frac{2\ \text{mol}\ \text{H}_2\text{O}}{2\ \text{mol}\ \text{H}_2}\right)\left(\frac{18\ \text{g}\ \text{H}_2\text{O}}{\text{mol}\ \text{H}_2\text{O}}\right) = 9.45\ \text{g}\ \text{H}_2\text{O}$

     Answer→ Oxygen is the limiting reagent; 1.37 grams of water

Percent Yield

Thursday, 14 February, 2019

- $\% \text{ Yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right)100$

1. When 32.97 grams of propane (C_3H_8) is burned in 15.1 liters of oxygen at STP, the percent yield of carbon dioxide is 92.1%. Calculate the mass of carbon dioxide recovered.
    - $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

      $32.97 \text{ g } C_3H_8 \left(\frac{\text{mol } C_3H_8}{44.11 \text{ g } C_3H_8}\right)\left(\frac{3 \text{ mol } CO_2}{\text{mol } C_3H_8}\right)\left(\frac{44.01 \text{ g } CO_2}{\text{mol } CO_2}\right) = 98.6858 \text{ g } CO_2$

      $15.1 \text{ L } O_2 \left(\frac{\text{mol } O_2}{22.4 \text{ L } O_2}\right)\left(\frac{2 \text{ mol } CO_2}{5 \text{ mol } O_2}\right)\left(\frac{44.01 \text{ g } CO_2}{\text{mol } CO_2}\right) = 17.8005 \text{ g } CO_2$

      Limiting reactant: O_2

      $(17.8005 \text{ g } CO_2)0.921 = 16.3942 \text{ g } CO_2$

   Answer→ 16.4 grams of carbon dioxide recovered
2. If 80.2 grams of magnesium reacts with excess silver nitrate to produce 98.64 grams of silver, what is the percent yield?
    - $Mg + 2AgNO_3 \rightarrow Mg(NO_3)_2 + 2Ag$

      $80.2 \text{ g Mg} \left(\frac{\text{mol Mg}}{24.31 \text{ g Mg}}\right)\left(\frac{2 \text{ mol Ag}}{\text{mol Mg}}\right)\left(\frac{107.87 \text{ g Ag}}{\text{mol Ag}}\right) = 711.7379 \text{ g Ag}$

      $\frac{98.64 \text{ g Ag}}{711.7379 \text{ g Ag}} = 0.138590$

   Answer→ 13.86%

Gases

Friday, 8 March, 2019

- No definitive shape or volume
- Kinetic-Molecular Theory
  - $E_K = \frac{mv^2}{2}$
  - Assumptions
    1. Gas molecules are small and spread far apart.
    2. Gas molecules are in constant motion.
    3. Gas molecules collide with each other and containers.
    4. Gas molecule collisions are perfectly elastic.
    5. Intermolecular forces between gas molecules are essentially negligible.
- Variables
  - **Pressure (*P*)**- Force over area; measured in Pascals (Pa)
    - 1 atm = 101325 Pa = 760 mmHg = t60 torr = 14.1 psi $\left(\frac{\text{lb}}{\text{in}^2}\right)$
  - **Volume (*V*)**- Space occupied by matter; measured in liters (L)
  - **Quantities (*n*)**- Amount of gas; measured in moles (mol)
  - **Temperature (*T*)**- Measure of average kinetic energy; measured in Kelvin (K)
- Gas Laws
  - Boyle's Law
    - Temperature is held constant.
    - There is an inverse relationship between pressure and volume.
    - $P_1V_1 = P_2V_2$
  - Charles's Law
    - Pressure is held constant.
    - There is a direct relationship between volume and temperature.
    - $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
      - $V_1T_2 = V_2T_1$
  - Gay-Lussac's Law
    - Volume is held constant.
    - There is a direct relationship between pressure and temperature.
    - $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
      - $P_1T_2 = P_2T_1$
  - Combined Gas Law
    - $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
  - Dalton's Law of Partial Pressures
    - $P_T = P_1 + P_2 + P_3 + \cdots$
  - Ideal Gas Law
    - Changes at low temperature or high pressure
    - $PV = nRT$
    - $R = 0.08205 \frac{L * atm}{mol * K}$
  - Graham's Law of Effusion
    - Ratio of diffusion rates

- $$\frac{\text{rate of gas A}}{\text{rate of gas B}} = \sqrt{\frac{\text{molar mass of gas A}}{\text{molar mass of gas B}}}$$

Solutions

Thursday, 28 March, 2019

- Homogenous Mixture
  - **Solute**- The part that dissolves
  - **Solvent**- The part that does the dissolving
  - "Like dissolves like."
    - Polar dissolves polar, and nonpolar dissolves nonpolar.
- $2\mathrm{AgNO_3}(aq) + \mathrm{MgCl_2}(aq) \rightarrow \mathrm{AgCl}(s) + \mathrm{Mg(NO_3)_2}(aq)$
  - $\mathrm{Ag^+}(aq) + \mathrm{Cl^-}(aq) \rightarrow \mathrm{AgCl}(s)$
- Molarity (M)
  - $\mathrm{M} = \frac{\mathrm{mol_{solute}}}{\mathrm{L_{solution}}}$
  - If you need to make a 375 mL of 1.2 M of sodium hydroxide, what do you need?

    $$\frac{1.2\ \mathrm{mol\ NaOH}}{\mathrm{L}\left(\frac{1000\ \mathrm{ml}}{\mathrm{L}}\right)} = \frac{x\ \mathrm{g\ NaOH}\left(\frac{\mathrm{mol}}{40\ \mathrm{g}}\right)}{375\ \mathrm{mL}}$$

    $$18\ \mathrm{g\ NaOH} = x$$

    Answer→ 18 grams sodium hydroxide
- **Dilutions**- Adding solvent to reduce concentrations
  - $M_1V_1 = M_2V_2$
- Molality (m)
  - $m = \frac{\mathrm{mol_{solute}}}{\mathrm{kg_{solvent}}}$
  - Calculate molality of 422 grams of sodium chloride diluted in 475 mL H_2O.

    $$\frac{422\ \mathrm{g\ NaCl}}{475\ \mathrm{mL\ H_2O}}\left(\frac{\mathrm{mol\ NaCl}}{58.44\ \mathrm{g\ NaCl}}\right)\left(\frac{\mathrm{mL\ H_2O}}{\mathrm{g\ H_2O}}\right)\left(\frac{1000\ \mathrm{g\ H_2O}}{\mathrm{kg\ H_2O}}\right) = 15.2023\ m$$

    Answer→ 15.2 molal sodium chloride
- **Unsaturated**- If there is less solute than maximum amount
- **Saturated**- If there is more solute
- **Supersaturated**- If there is more solute than the maximum amount
- Factors Affecting Solubility
  - Nature of solute and solvent
    - Polar and nonpolar
  - Temperature
    - ↑ Temperature = ↑ Solubility
  - Pressure
    - For gases
    - ↑ Pressure = ↑ Solubility
- **Colligative Processes**- A property dependent on the concentration of solute added but independent of the nature of the solute
  - **Vapor Pressure Reduction**- Adding nonvolatile solute lowers vapor pressure
  - **Freezing Point Depression**- Adding nonvolatile solute lowers the freezing point
    - $\Delta T_f = K_f m i$
      - ΔT_f: Change in freezing point
      - K_f: Freezing point constant

            - $-1.86\frac{°C}{m}$ for water
    - Calculate the change in freezing point when 325 grams $C_6H_{12}O_6$ are added to 3.259 kilograms water.
        - $\frac{325\text{ g } C_6H_{12}O_6}{3.259\text{ kg } H_2O}\left(\frac{\text{mol } C_6H_{12}O_6}{180\text{ g } C_6H_{12}O_6}\right) = 0.5540\ m$
          $\Delta T_f = \left(-1.86\frac{°C}{m}\right)(554.0213\ m)(1)$
          $\Delta T_f = -1.03°C$
      Answer→ –1.03°C
    - Calculate the freezing point depression for 62 g NaCl in 9876 g H_2O.
        - $\frac{62\ g\ NaCl}{9876\ g\ H_2O}\left(\frac{mol\ NaCl}{58.44\ g\ NaCl}\right)\left(\frac{1000\ g\ H_2O}{kg\ H_2O}\right) = 0.1074\ m$
          $\Delta T_f = \left(-1.86\frac{°C}{m}\right)(0.1074\ m)(2)$
          $\Delta T_f = -0.3996°C$
      Answer→ –0.40°C
- **<u>Boiling Point Elevation</u>**- Adding nonvolatile solute increases the boiling point
    - $\Delta T_b = k_b m i$
        - ΔT_b: Change in boiling point
        - k_b: Boiling point constant
            - $0.512\frac{°C}{m}$ for water
    - What is the new boiling point for a solution containing 14 g $MgCl_2$ in 3509 grams of water?
        - $\frac{14\text{ g } MgCl_2}{3509\text{ g } H_2O}\left(\frac{\text{mol}}{95.21\text{ g}}\right)\left(\frac{1000\text{ g}}{\text{kg}}\right) = 0.0419\ m$
          $\Delta T_b = \left(0.512\frac{°C}{m}\right)(0.0419\ m)(3)$
          $\Delta T_b = 0.0644°C$
      Answer→ 100.06°C
- **<u>Osmotic Pressure</u>**- Pressure required to prevent osmosis

Intermolecular Forces

Monday, 8 April, 2019

- **Intramolecular Forces**- Bonding (within molecules)
- **Intermolecular Forces**- Between molecules or atoms; force of attraction between neighboring particles
  - Not as strong as bonds
- Types
  - **London Dispersion Forces**- Force of attraction between temporary dipoles
    - **Dipole**- Polar molecule
    - Exist between *all* atoms and molecules
    - Only intermolecular force for nobles gases and nonpolar molecules
    - If two temporary dipoles are near each other, there will be a force of attraction between the δ^- and the δ^+ side of the atom/molecule.
    - Strength increases with increasing size.
      - The bigger the atom/molecule, the greater the force.
        - Due to polarizability
          - **Polarizability**- Ability to be polar
      - Oganesson has the highest boiling point.
  - **Dipole-Diploe**- Force of attraction between two diploes
    - Will also have London Dispersion Forces
    - *i.e. between molecules of hydrogen sulfide*
    - Stronger than London Dispersion Forces
    - Between true dipoles, not temporary dipoles
    - **Hydrogen Bonding**- Between a polarized hydrogen and a nitrogen, oxygen, or fluorine on another molecule
      - Nitrogen, oxygen, and fluorine are the most electronegative.
      - Not actually a bond, just a strong intermolecular force
      - Special case of dipole-dipole
        - Stronger than regular dipole-dipole
      - *i.e. between water molecules; between base pairs in DNA*
  - **Ion-Dipole**- Force of attraction between an ion and a dipole
    - *i.e. between NaCl and H_2O (Na^+ and Cl^-)*

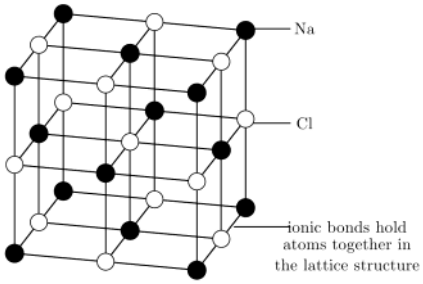
Properties of Liquids

Monday, 8 April, 2019

- Liquids are more ordered than gases.
  - Stronger intermolecular forces
  - Lower mobility of particles
- Gases and liquids are fluids due to their mobility.
- Liquids have a relatively high density (mass per volume).
  - Hundreds of times more dense than the gaseous state.
  - Ten percent less dense than the solid state.
- Liquids are much more compressible than gases because their particles are closer together,
- Liquids, like gases, diffuse.
  - High to low
  - Not as fast
  - If temperature increases, rate of diffusion increases.
- **Surface Tension**- A force that tends to pull adjacent parts of a liquid's surface together, which decreases the surface area
- **Capillary Action**- The attraction of the surface if a liquid to the surface of a solid; pulls liquids upwards against the force of gravity
- **Viscosity**- Resistance to flow

Solids

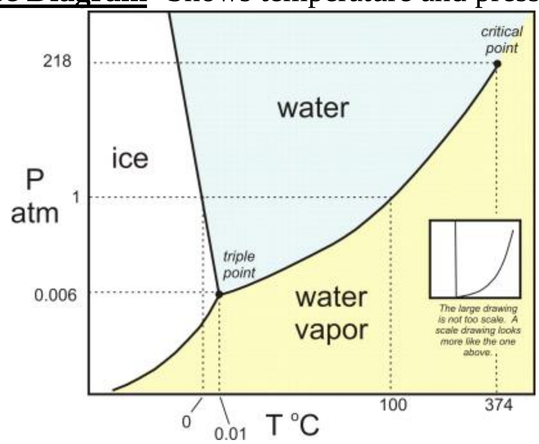
Friday, 12 April, 2019

- **Crystalline**- Regular arrangement of particles; have a regular repeating pattern of units in a lattice
  - *i.e. NaCl*
  - 

  - Types

	Ionic	**Molecular**	**Covalent Network**	**Metallic**
Parts of Crystal	Cations and Anions	Individual molecules	Atoms covalently bonds to each other	Metals, alloys
Strongest force in binding crystal	Ionic bond	Intermolecular forces	Covalent bond	Metallic bond
Hardness	Hard	Soft	Very hard	Malleable
Brittleness	Brittle	Crumbly	Very brittle	Malleable
Melting point	High	Low	Very high	Variable
Conductivity	No (unless molten)	No	No (insulators)	Yes
Solubility	Usually soluble in H_2O	Polar dissolves polar Nonpolar dissolves polar	Insoluble	Insoluble
Example	NaCl, $Ca(NO_2)_2$	Ice (H_2O), Dry ice (CO_2)	Diamond	Iron, gold

- **Amorphous Solid**- Irregular arrangement of particles; no regular pattern
  - "Supercooled liquids"
  - *i.e. plastics, rubber, glass*
- Changes of State
  - Require energy
    - Going to a less condensed state is an endothermic reaction.
  - Terminology
    - Solid → Liquid: Melting
    - Liquid → Gas: Vaporizing
    - Liquid → Solid: Freezing
    - Gas → Liquid: Condensation

        - Solid → Gas: Sublimation
        - Gas → Solid: Deposition
    - **<u>Phase Diagram</u>**- Shows temperature and pressures for different phases
        - 

        - The line is the equilibrium point.
        - At the triple point, all three phases exist at once.
        - Once you have passed the critical point, if will always be a gas despite pressure or temperature.
    - Special Phases of Water
        - Less dense as a solid than a liquid
        - Relatively high melting and boiling point
        - High surface tension
        - High heat of vaporization
        - High specific heat capacity
        - Universal solvent

Acids/Bases

Tuesday, 16 April, 2019

Acids	Bases

• Taste sour
  • *i.e. citrus*
• React with some metals to produce H_2
• Change color of acid/base indicators
• Conduct electricity in solution
	• Taste bitter
	  • *i.e. soap*
	• Feel slippery
	• Change color of acid/base indicators
	• Conduct electricity in solution

- Definitions
	Acid	Base
Arrhenius	Generates H^+ in solution *i.e. HCl,* CH_3COOH	Generate OH^- in solution
Bronsted-Lowry	Proton (H^+) donor	Proton (H^+) acceptor Bases do not need to contain a hydroxyl group *i.e.* NH_3*, amine groups*
Lewis	Electron pair acceptor	Electron pair donor
  - **Amphoteric**- Can act as an acid or a base
    - *i.e.* H_2O *can accept a proton* (H_3O^+) *or donate a proton* (H^+ *and* OH^-).
      - $\mathrm{H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-}$
        - H_3O^+ becomes the conjugate acid.
        - OH^- becomes the conjugate base.
          - Conjugates will be on the right side of the equation.
            - *i.e.* $HC_2H_3O_2 + H_2O \rightleftharpoons C_2H_3O_2^- + H_3O^+$
              - *Acetic acid is the acid, and water is the base. The conjugate acid is acetate, and the conjugate base is hydronium.*

Concentrations of Acids and Bases

Monday, 22 April, 2019

- $\mathrm{H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-}$
    - Autoionization of H_2O
    - At 25 °C, $K_w = 1.0 \times 10^{-14} = \left[\mathrm{H_3O^+}\right][\mathrm{OH^-}]$
- *A solution contains 2.65 × $10^{\ 4}$ M OH . Calculate [H_3O^+].*
    - $1.0 \times 10^{-14}\mathrm{M} = \left[\mathrm{H_3O^+}\right][\mathrm{OH^-}]$
      $1.0 \times 10^{-14}\mathrm{M} = \left[\mathrm{H_3O^+}\right]\left(2.65 \times 10^{-4}\mathrm{M}\right)$
      $3.8 \times 10^{-11}\mathrm{M} = \left[\mathrm{H_3O^+}\right]$

  Answer→ 3.8×10^{-11} M
- Strong Acids and Bases
    - Strong acids or base
        - Dissociates fully in solution
        - Strong Acids
            1) Halogens
                - HCl, HBr, HI
                - Not HF
                    - Fluorine is more electronegative than the oxygen in the water.
            2) Perchloric and chloric acid
                - $HClO_4$ and $HClO_3$
            3) Nitric and sulfuric acid
                - HNO_3 and H_2SO_4
        - Strong Bases
            - Alkali Metals
                - LiOH, NaOH, KOH, RbOH
            - Alkali Earth Metals
                - $Ca(OH)_2$
                - $Sr(OH)_2$
                - $Ba(OH)_2$
    - Weak acid or base
        - Partially dissociates in solution
        - *i.e. acetic acid*
- **<u>Monoprotic</u>**- One proton (hydrogen ion)
    - *i.e. $HC_2H_3O_2$ and HCl are monoprotic acids.*
- **<u>Polyprotic</u>**- More than one proton (hydrogen ion)
    - *i.e. H_2SO_4 and H_2SO_3 are polyprotic acids.*
- *Practice*
    a. If you have 2.65×10^{-4} M $Ca(OH)_2$, what is your concentration of hydroxide?
        - $[\mathrm{OH^-}] = 5.30 \times 10^{-4}\mathrm{M}$
    b. A solution containing 1.23×10^{-4} M HCl has what concentration of OH^-?
        - $[\mathrm{OH^-}] = 8.1 \times 10^{-11}\mathrm{M}$
    c. A solution contain 8.62×10^{-10} M $Sr(OH)_2$. Calculate [H^+].
        - $\left[\mathrm{H^+}\right] = 5.8 \times 10^{-6}\mathrm{M}$
- pH Scale
    - Measures the hydrogen ion concentration
    - $\mathrm{pH} = -\log[\mathrm{H^+}]$
    - 7 is neutral.

- Below 7 is acidic.
- Above 7 is basic/alkaline.

- pOH Scales
  - Measure the hydroxide ion concentration
  - $\mathrm{pH} = -\log[\mathrm{OH}^-]$
  - 7 is neutral.
  - Below 7 is basic/alkaline.
  - Above 7 is acidic.
- $\mathrm{pH} + \mathrm{pOH} = 14.00$

Titrations

Thursday, 25 April, 2019

- **Titration**- Controlled addition and measurement of the amount of a solution of known concentration to a solution of unknown concentration, resulting in a complete reaction
- Acid-Base Neutralization Reaction
    - Acid + Base ⇌ Salt + H_2O
        - *i.e.* $HCl + NaOH \rightleftharpoons NaCl + H_2O$
    - You need equal amounts of acids (H^+) and bases (OH^-) to neutralize.
    - Equivalence point
        - $\text{mol H}^+ = \text{mol OH}^-$
        - Strong Acid — Strong Base: 7
- Done with a burette
    - Titrant in the burette and an unknown concentration in the beaker/flask
    - Allows you to control and measure the amount of added titrant
    - How to know when the titration reaction is done:
        - Using a pH meter or a pH probe
        - Acid-base/pH indicator
            - Changes color in certain pH range
            - **Endpoint**- When indicator changes color
- *Practice*
    a. In a titration, 27.5 mL of 0.0154 M $Ba(OH)_2$ is added to a 20.0 mL sample of HCl of unknown concentration until the equivalence point is reached. What is the concentration of the HCl?
        - $M_A V_A = M_B V_B$
          $M_A(20.0\text{ mL}) = (0.0308\text{ M})(27.4\text{ mL})$
          $M_A = 0.042196\text{ M}$

    Answer→ 0.0422 molar hydrochloric acid
- Titration Curves
    - mL of base or acid is on the *x*-axis
    - pH is on the *y*-axis
    - **Buffering Region**- The point in which the line is changing gradually

Thermodynamics

Thursday, 9 May, 2019

- **Thermodynamics**- Study of transfers of energy as heat that accompany chemical reactions and physical changes
- **Temperature**- Measure of the average kinetic energy of a substance; measure in °C or K
    - $E_k = \frac{mv^2}{2}$
    - Heat SI unit = Joule ($\mathcal{J}$)
    - **Heat**- Energy transferred between 2 substances due to temperature differences
        - There's no such thing as "cold." It is caused by the absence of heat.
- **Specific Heat**- Energy required to raise 1 gram of a substance by 1 degree
    - $\mathrm{H_2O} = 4.184 \frac{\mathcal{J}}{\mathrm{g} * °\mathrm{C}}$
    - $q = mC\Delta T$
        - $q =$ heat ($\mathcal{J}$)
        - $m =$ mass (g)
        - $\Delta T = T_f - T_i$
- Energy absorbed/released by a reaction is represented by ΔH at a constant pressure.
    - $\Delta H =$ Enthalpy change
    - $2\mathrm{S} + 3\mathrm{O_2} \rightarrow 2\mathrm{SO_3}\ \Delta\mathrm{H} = -791.4\mathrm{k}\mathcal{J}$
    - $2\mathrm{S} + 3\mathrm{O_2} \rightarrow 2\mathrm{SO_3} + 791.4\mathrm{k}\mathcal{J}$
    - *If 6.44 g of sulfur reacts with excess O_2, how many kilojoules of heat are released?*
        - $6.44\ \mathrm{g\ S}\left(\frac{\mathrm{mol\ S}}{32.06\ \mathrm{g\ S}}\right)\left(\frac{-791.4\ \mathrm{k}\mathcal{J}}{2\ \mathrm{mol\ S}}\right) = -79.5\mathrm{k}\mathcal{J}$
          Answer→ –79.5 kilojoules

Hess' Law

Monday, 13 May, 2019

- **Hess' Law**- Overall enthalpy change for a reaction is equal to the sum of the individual enthalpy changes for the individual steps of a process
- *Examples*
  a. Calculate ΔH for $PCl_3 + Cl_2 \rightarrow PCl_5$
     - $2P + 3Cl_2 \rightarrow 2PCl_3 \quad \Delta H = -640\ k\mathcal{J}$
     - $2P + 5Cl_2 \rightarrow 2PCl_5 \quad \Delta H = -886\ k\mathcal{J}$
     - $PCl_3 \rightarrow P + 1.5Cl_2 \quad \Delta H = +320\ k\mathcal{J}$
       $P + 2.5Cl_2 \rightarrow PCl_5 \quad \Delta H = -443\ k\mathcal{J}$
       $PCl_3 + Cl_2 \rightarrow PCl_5 \quad \Delta H = -123\ k\mathcal{J}$

     Answer→ $\Delta H = -123\ k\mathcal{J}$
  b. Calculate ΔH for $2F_2 + 2H_2O \rightarrow 4HF + O_2$
     - $H_2 + F_2 \rightarrow 2HF \quad \Delta H = -542.2\ k\mathcal{J}$
     - $2H_2 + O_2 \rightarrow 2H_2O \quad \Delta H = -571.6\ k\mathcal{J}$
     - $2H_2 + 2F_2 \rightarrow 4HF \quad \Delta H = -1084.41\ k\mathcal{J}$
       $2H_2O \rightarrow 2H_2 + O_2 \quad \Delta H = 571.6\ k\mathcal{J}$
       $2H_2O + 2F_2 \rightarrow 4HF + O_2 \quad \Delta H = -512.8\ k\mathcal{J}$

     Answer→ $\Delta H = -512.8\ k\mathcal{J}$

Calorimetry

Wednesday, 15 May, 2019

- Calorimeter
- $q = mC\Delta T$
- *Examples*
  a. When a 12.8 gram sample of KCl dissolves in 75.0 grams of water, the temperature drops from 31.0 °C to 21.6 °C. Calculate ΔH for the process $\mathrm{KCl} \rightarrow \mathrm{K}^{+} + \mathrm{Cl}^{-}$.
     - $q_{\mathrm{surr}} = mC\Delta T$

       $$= (75.0\ \mathrm{g})\left(4.184\,\frac{\mathcal{J}}{\mathrm{g} * {}^{\circ}\mathrm{C}}\right)(21.6\ {}^{\circ}\mathrm{C} - 31.0\ {}^{\circ}\mathrm{C})$$

       $$= -2949.72\ \mathcal{J}$$
     - $q_{\mathrm{rxn}} = -q_{\mathrm{surr}}$

       $$= 2949.72\ \mathcal{J}$$
     - $\Delta H = \dfrac{q_{\mathrm{rxn}}}{\mathrm{mol}}$

       $$= \frac{2949.72\ \mathcal{J}}{0.172\ \mathrm{mol}}$$

       $$= 17200\,\frac{\mathcal{J}}{\mathrm{mol}}$$

     Answer→ 17,200 joules per mole
  b. What is the specific heat of silicon if the temperature of a 4.11 gram sample is increase by 3.8 °C when 11.1 joules of heat is added.
     - $q = mC\Delta T$

       $(11.1\ \mathcal{J}) = (4.11\ \mathrm{g})C(3.8\ {}^{\circ}\mathrm{C})$

       $C = 0.71\,\dfrac{\mathcal{J}}{\mathrm{g} * {}^{\circ}\mathrm{C}}$

     Answer→ $0.71\,\frac{\mathcal{J}}{\mathrm{g}*{}^{\circ}\mathrm{C}}$