

CHEMISTRY CONTENT FACTS

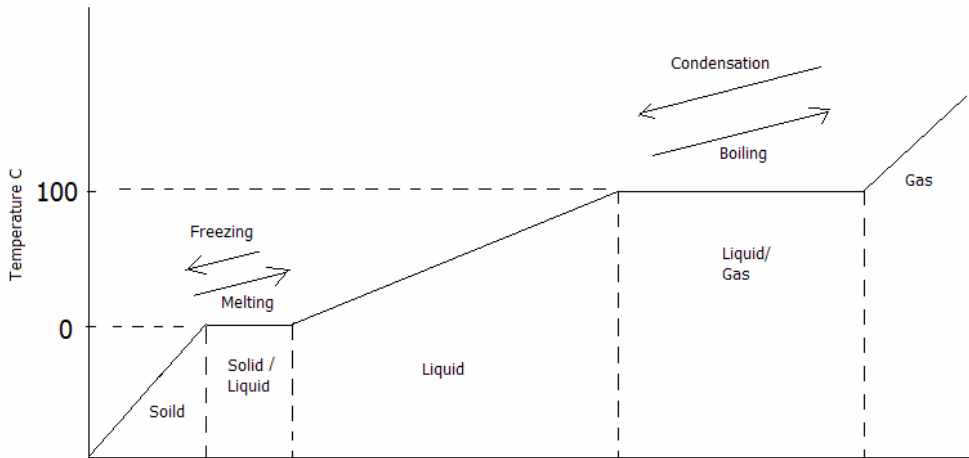
The following is a list of facts related to the course of Chemistry. A deep foundation of factual knowledge is important; however, students need to understand facts and ideas in the context of the conceptual framework. This list is not intended to provide a comprehensive review for State and National Assessments. Its purpose is to provide a highlight of the factual material covered in Chemistry. This list is not all inclusive, be sure to check Nevada State Standards and your district syllabi.

Matter and Energy

- Properties of **solids** - definite shape & volume; regular geometric pattern; crystalline structure
- Properties of **liquids** - no definite shape, but definite volume
- Properties of **gases** - no definite shape or volume, random particle motion
- **Elements** - all atoms have the same ATOMIC #. Can **NOT** be broken down chemically
- **Mixture** - 2 or more elements physically combined. There are different types of mixtures
 - heterogeneous (unevenly mixed – chocolate chip ice cream)
 - homogeneous (evenly mixed SOLUTION - clear tea)
- **Physical change** - no change in the identity of the substance (i.e. gas to liquid to solid, or cutting, crushing, etc.)
- **Chemical change** – substance(s) changes into new substance(s) with NEW properties ($H_2 + O_2 \rightarrow H_2O$)
- **Law of Conservation of Mass:** In a chemical reaction, mass remains constant
- For heat flow problems: know the formula $\Delta q = m(\Delta t)C_p$
- q = heat flow, m = mass, Δt = change in temperature, C_p = specific heat
- Temperature is the measure of the average kinetic energy of particles *
- Convert from Celsius to Kelvin (+ 273) and back (- 273) *
- The potential energy of a system is considered to be the HEAT of the system
- Fixed points on a thermometer - $0^\circ C$ - freezing/melting point of H_2O ; $100^\circ C$ boiling/condensation point of H_2O ; at standard pressure (2 points needed to create a thermometer) *
- Gas law problems - Combined gas law: (Temp must be in Kelvin) $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- Boyle's Law – Constant Temperature, Pressure and Volume vary inversely $P_1V_1 = P_2V_2$
- Charles's Law – Constant Pressure, Volume and Temperature are directly proportional
$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } \frac{V_1}{V_2} = \frac{T_1}{T_2}$$
- **STP** - Standard Temperature and Pressure of $0^\circ C$ or 273 K and 101.3kPa or 760 mm Hg or 1 atm
- **Density** is directly proportional to mass and inversely proportional to volume $d = \frac{m}{V}$

- **Sublimation** a substance turns directly from a solid to a gas ex. $\text{CO}_2(\text{s}) \rightarrow \text{CO}_2(\text{g})$; $\text{I}_2(\text{s})$ (purple crystals) $\rightarrow \text{I}_2(\text{g})$ (purple gas)
- **Phase change diagrams:** Melting/Boiling and Freezing/Condensation [See diagram]

Heating / Cooling Curve for Water



Heat

- **Kinetic molecular theory**
- all matter is composed of very small particles called atoms
- atoms are in constant, random motion
- collisions between atoms are elastic
- Ideal gases are point masses
- **Ideal Gas Law:** $PV = nRT$ P = pressure (atm or kPa), V = volume (L or dm^3), n = # moles, T = temperature (K), R = 0.0821 L·atm/mol·K or 8.31 dm^3 ·kPa/mol·K
- Real gases **BEHAVE LIKE IDEAL GASES** except at high Pressure and/or low Temperature
 - When there **IS** an attraction between particles (van der Waals forces) and
 - The volume of particles is **NOT** negligible, since atoms/molecules are close together
- *****HYDROGEN and HELIUM** are the most **IDEAL** gases. Also, noble gases act very ideally.
 - THE SMALLER THE MOLECULES ARE THE MORE IDEALLY THEY BEHAVE
- **Molar Heat of fusion** - the amount of heat needed to melt one mole of a solid; for H_2O it is 6.0 kJ/mol
- **Molar Heat of vaporization** - the amount of heat needed to vaporize one mole of a solid; for H_2O it is 40.7 kJ/mol
- **Boiling point** - the temperature at which the vapor pressure of a liquid = the atmospheric pressure. The normal boiling point when the atmospheric pressure = 760 mm Hg = 1 atm = 101.3 kPa
- **Vapor pressure** - depends on;

- a) the temperature of the liquid
- b) the strength of intermolecular forces (i.e. the stronger the van der Waals forces the stronger the Intermolecular forces are and the lower the vapor pressure is)
- c) **Law of partial pressures-Dalton's Law:** the sum of all the partial pressures in a mixture of gases is equal to its total pressure $P_{total} = P_1 + P_2 + P_3$

Atomic Structure

- **Parts of the atom:**
 - Proton: (+) charged; 1 atomic mass unit
 - Neutron: not charged; ~1 atomic mass unit
 - Electron: (-) charged; 1/1836 atomic mass unit
- **Nucleon** - particles found in the nucleus (protons & neutrons)
- **Nucleus** - contains most of the mass of the atom; has a positive charge; The # of protons is called the atomic number and produces the nuclear charge
- 1 U (formerly amu) - the atomic mass unit; based on 1/12 the mass of a **carbon 12** atom
- In a neutral atom the # of protons = the # of electrons
- **Atomic #** - the # of protons in an atom; used to identify the element
- **Atomic mass** = the # of protons + the # of neutrons
- **Isotopes** - elements that have the same atomic # but different atomic masses due to a difference in the # of neutrons in the nucleus
- To figure out the # of neutrons in an element subtract the atomic # **FROM** the mass number
- ${}_6\text{C}^{14}$ has 6 protons, 6 electrons and 8 neutrons
- Atomic mass is really a **weighted average** of all of the isotopes that exist in nature for that element. i.e. Carbon's atomic mass = 12.011 because there is ${}_6\text{C}^{12}$ and ${}_6\text{C}^{14}$ in nature but ${}_6\text{C}^{12}$ is more abundant and therefore skews the average toward 12
- **Quantized model of the atom** - states that electrons are most probably in certain energy levels. An absorption of energy will cause electrons to **TEMPORARILY** jump to higher levels and when the electrons fall back down to lower levels they **EMIT** this energy in the form of light
- 13. **Valence electrons** - electrons in the outermost energy levels. i.e. ${}_9\text{F}^{19} 1s^2 2s^2 2p^5$ has 7 valence electrons **SINCE** the outer most principle energy level is the 2nd one. Kernel electrons are the electrons **NOT** considered to be part of the valence shell
- **Electron dot diagram** - uses dots for the valence electrons
- **Orbital diagrams:** illustrate the orbitals occupied by 0, 1, or 2 electrons. Arrows represent the electrons & two electrons or arrows can fit into each orbital. The electrons in the same orbital **MUST** spin in opposite directions

- **Hund's rule** - before an orbital can get a second electron each orbital in that sublevel must have at least one electron
- **Order of filling sublevels:** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$: **WHY?** The $4s^2$ sublevel needs less energy to fill than the $3d^{10}$ sublevel

Bonding

- When a bond is formed energy is released (exothermic); when a bond is broken energy is absorbed (endothermic)
- Atoms bond together to form **OCTETS** (this is lower in energy than unbonded atoms)
- Metals tend to form positive ions by loss of electrons
(these cations are smaller than the neutral atoms: Ionic radii < than atomic radii)
- Nonmetals tend to form negative ions by gaining electrons (anions formed are larger than the neutral atoms: Ionic radii > atomic radii)
- **A chemical bond** - results from the simultaneous attraction of electrons by two nuclei
- **Ionic bonds** - formed between metal and nonmetal; created by a transfer of electrons; electronegativity difference ≥ 1.7
- **Covalent bond** - formed by the sharing of electrons; electronegativity difference < 1.7
- **Electronegativity** - the affinity for electrons in a covalent bond. Highest: **Fluorine 4.0**
- Diatomic molecules are considered to have **NONPOLAR** covalent bonding
- *Helium & Hydrogen* need only 2 electrons to fill outer shell. The first energy level can contain a maximum of two electrons
- **Coordinate covalent bonds** - a covalent bond where both of the shared electrons are donated by one of the atoms. Usually found in polyatomic ions
- **Ions:** K^+ and Cl^- have the same # of electron (18) since formation of ions are caused by the loss or gaining of ELECTRONS
- **Ionization energy:** the amount of energy required to remove the outermost electron from an atom
- **Ionic solids:** high melting & boiling point; hard; do not conduct electricity UNLESS dissolved in water and/or in molten form
- **Metallic solids:** mobile electrons, conductors in solids phase, malleable, ductile, only metal that is a liquid at room temperature is Hg (mercury)
- **Molecular solids:** molecules held together by van der Waals forces; low melting & boiling points; poor conductors; are soft. ex. Sugar $C_6H_{12}O_6$

- **Network solids:** held together by a network of covalent bonds; high melting & boiling points. Extremely poor conductors of heat & electricity. i.e. SiO₂, diamond - tetrahedral bonding (C_n), graphite (C_n) - hexagonal bonding
- Van der Waals forces - attractive forces that exist between ALL particles. They increase when particles
 - Increase in size
 - Get closer together
- **Hydrogen bonds** - attractive for between molecules that contain hydrogen and atoms of small atomic radius and **HIGH ELECTRONEGATIVITIES**. i.e. H₂O and HF. These bonds result in some compounds having higher boiling points than expected
- **Polar molecules** - molecules in which there is a localization of charge that causes part of the molecule to be slightly positively charged [δ^+] and part of the molecule to be negatively charged [δ^-]. Bonded electrons are shared unequally (polar bonds) and these polar bonds are asymmetrically arranged
- **Nonpolar molecule** - there may still be localization of charge but there is no NET dipole of electrons in any particular direction.

Periodic Table

- **Periodic law** - states that the physical and chemical properties of the elements are a periodic function of their atomic numbers and results in the element arrangement on the periodic table
- **Elements are classified in 3 categories**
 - *Metals* - left of stairs
 - *Nonmetals* - right of the stairs
 - *Metalloids* - touching the stairs
- **Trends** - as you go from left to right across the table in a period
 - Metallic character decreases
 - Atomic radius decreases
 - Ionization energy increases
 - Electronegativity increases
- As you go **down** a group
 - Metallic character increases
 - Atomic radius increases
 - Ionization energy decreases
- **Metalloids** - have both metal and nonmetal properties. Contact the "staircase".
- **Group 1 metal** - alkali metals; form 1⁺ ions
- **Group 2** - alkaline earth metals; form 2⁺ ions

- **Group 18 noble gases**; generally non-reactive. Kr and Xe can form some bonds
- **Group 17 halogens** - contain elements in ALL three phases. F₂ & Cl₂ are gases, Br₂ is a liquid and I₂ is a solid
- Elements in the same period fill up the **SAME** principle energy levels
- Elements in the same groups have the same # of valence electrons
- The most active metals are in the lower left corner.
- The most active nonmetals are in the upper right corner (excluding noble gases)
- The **MOST** active elements for the MOST stable compounds! i.e. RbF
- Monatomic molecules (one atom) He, Ne, Ar, Kr, Rn
- Diatomic elements (two atoms) H₂, N₂, O₂, F₂, Cl₂, Br₂, I₂
- Transition elements -
 - Generally produce **COLORED SOLUTIONS**
 - found in the middle of periodic table
 - emit color in flame test as electrons fall back DOWN from the excited state
 - lose both s & d electrons & therefore have multiple oxidation states
- Van der Waals forces increase as you go down a group since the size of the atom increase. This causes the boiling and melting points to increase as well. Remember this when you get to ORGANIC chemistry
- Atomic radius decreases as you go across a period since there is an increase of effective nuclear charge (# of protons) which pulls the electrons in same energy level closer thereby decreasing the size of the atom

Stoichiometry and Mathematics

- **Mole** of any gas occupies 22.4 dm³ at STP and contains 6.02 X 10²³ molecules
- $\# \text{ of moles} = \frac{\text{given mass (grams)}}{\text{molar mass (add up masses from periodic table)}}$
- **Avogadro's Law** - equal volumes of gases at same temperature and pressure contain equal # of molecules
- **Molarity** = $\frac{\# \text{ of moles of solute}}{\text{Liter of solution}}$
- **Molality** = $\frac{\# \text{ of moles of solute}}{\text{Kg solvent}}$
- **Solubility**: like dissolves like (polarities)
- **Unsaturated solution** - holds less solute than the maximum for given temperature and pressure

- **Saturated** - holds the exact amount of solute the solvent can hold for given temperature and pressure
- **Super saturated** - holds more than the maximum amount of solute for that temperature and pressure
- **Concentrated solution** - holds a large amount of solute for volume solvent
- **Dilute solution** - holds a small amount of solute for volume solvent
- **Solubility of a solid** - (ability to dissolve) generally increases as temperature increases
- **Solubility of a gas** – generally increase as temperature decreases and pressure increases. Think of when soda goes flat (CO₂ escapes)
- **Boiling point elevation** - for every mole of particles dissolved in solution the boiling point increase by a constant amount. i.e. for H₂O, $k_b = 0.52 \text{ }^\circ\text{C/molal}$
- **Freezing point depression** - for every mole of particles dissolved in solution the freezing point decreases by a constant amount. i.e. for H₂O, $k_f = 1.86 \text{ }^\circ\text{C/molal}$
- When figuring out boiling point elevation and freezing point depression keep in mind that electrolytes (molecules that split into ions) create more moles in solution than the moles solute dissolved
- How do you know when a substance is an electrolyte? If it is *ionically* bonded it is an electrolyte. i.e. NaCl (salt) or NaOH (base) or it is polar covalently bonded and ionizes in H₂O (ex. HCl, HBr)
- **Molecular formula** - the actual # of atoms in the covalently bonded molecule. i.e. C₆H₁₂O
- **Empirical formula** - shows the simplest ratio of atoms in a substance i.e. C₆H₁₂O₆ → CH₂O
- Finding the empirical formula from percentages
 - Divide the percentages by the atomic masses (see periodic tables)
 - Divide the resulting numbers by the smallest result and this gives you your ratio for the empirical formula
 - Put this into a whole number ratio if fraction results
- Finding the molecular formula from percentages. You **MUST** be given the molar mass to do this
 - Divide the percentages by the atomic masses (see periodic tables)
 - Divide the resulting numbers by the smallest result and this gives you your ratio for the empirical formula
 - Figure out what the empirical formulas mass is and see how many times it goes in to your total mass
 - Then multiply subscripts by this number
- **Percentage comp.** = $\frac{\text{Total mass of element in compound}}{\text{Total mass of compound}} \times 100$
- **Percent error** = $\frac{|\text{experimental value} - \text{true value}|}{\text{true value}} \times 100$

Kinetics and Equilibrium

- **Heat of reaction (ΔH)** - the difference between the potential energy of the products and the reactants
- **Exothermic reactions** release energy, (ΔH) < 0, products formed are MORE stable compounds than the reactants
- **Endothermic reactions** absorb energy, (ΔH) > 0, products formed are LESS stable compounds than the reactants
- **Factors effecting the reactions rate**
 - **Catalyst** - speeds up the reaction by reducing the activation energy needed for the reaction. A catalyst does NOT effect the heat of reaction or the potential energy of the products or the reactants
 - **Increasing the concentration of one of the reactants** usually increases the rate
 - **Increase in temperature** gives more effective collisions and increases the rate
 - **Increase in surface area** of solids increases the reaction rate
- **Entropy (S)**- the randomness of a system. If $\Delta S > 0$ then there is an increase in entropy or Randomness and if $\Delta S < 0$ then there is a decrease
- Order of increasing entropy: solids \rightarrow liquids \rightarrow gases
- **Gibb's equation** $\Delta G = \Delta H - T \Delta S$ states whether or not a reaction occurs spontaneously or not. If $\Delta G < 0$ the reaction will occur spontaneously and if $\Delta G > 0$ the reaction is non-spontaneous. When $\Delta G = 0$ the system is at equilibrium
- $$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad aA (g) + bB (g) \leftrightarrow cC (g) + dD (g)$$
- When K_{eq} is large that means that the reaction favors the products
- When K_{eq} is small that means that the reaction favors the reactants.
- Remember that the coefficients in front of the compounds become the exponents in the equilibrium constant equation for all (g) and (aq) substances
- Effects of stresses on systems at equilibrium (Le Chatelier's Law)
 - Increasing concentrations of reactant(s) shifts system to produce more products (right)
 - Increasing concentrations of product(s) shifts system to produce more reactants (left)
 - Decreasing concentrations of reactants shifts system to the left
 - Decreasing concentrations of products shifts system to the right
 - Increasing temperature favors endothermic reaction
 - Decreasing temperature favors exothermic reactions
 - Increasing pressure favors side with fewer moles of gas
 - Decreasing pressure favors side with more moles of gas

- Adding or removing solid has no effect
- Adding or removing catalyst has no net effect
- Solubility product equation - K_{sp} = Dissociated ions ONLY [Ions are charged particles; +/-]
- When K_{sp} is large that means the substance is soluble
- When K_{sp} is small that means that the substance has low solubility
- **Ionization constant for acids**- same as equilibrium constant but you use K_a instead
- When K_a is large that means that the reaction favors the ionization. This is a strong ACID
- When K_a is small that means that the reaction favors the non dissociated part of the equation. This is a weak ACID

Acids and Bases

- **Electrolyte** – a compound that conducts electricity when melted or dissolved in aqueous solution. Usually ionically or polarly covalently bonded
- **Non-electrolyte** - a compound that does not conduct electricity in solution or when melted, non-polar covalently bonded
- **Arrhenius theory of Acid** forms H^+ ion, as the ONLY + ion in water solution, **Base** forms OH^- ion
- **Bronsted-Lowry Theory** - **Acid** = proton donor (loses H^+), **Base** = proton acceptor (gains H^+)
- **Salt** - a metal combined with a nonmetal [ex. NaCl, Na is the metal & Cl is the nonmetal]
- **Organic compounds**- contain with C. i.e. $C_6H_{12}O_6$ - usually NOT electrolytes. Except organic acids [functional group $-COOH$]
- **Traits of Acids**
 - Turns blue litmus red
 - pH less than 7.0
 - Reacts with metals to form salt and H_2 gas
 - Taste sour
 - React with base to form salt and water (neutralization)
 - The more they ionize, the better they conduct electricity
 - Contain more H^+ (H_3O^+) than (OH^-)
- **Traits of bases**
 - Turns red litmus blue, pink in phenolphthalein
 - pH greater than 7.0
 - Reacts with acids - neutralization
 - Taste bitter
 - Feel slippery

- The more they dissociate or ionize, the better they conduct electricity
- They contain more OH^- than H^+ (H_3O^+)
- **Neutralization - Acid + Base** \rightarrow salt + water $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ (net reaction) for strong acid and base
- In neutralization, moles of acid and moles of base must be **equal**
- Formula for **titration (neutralization)**
 - $\text{ACID [Molarity]} \times [\text{liters}] = \text{BASES [Molarity]} \times [\text{liters}]$
- List of conjugate acid-base pairs - Strongest acid = largest K_a and Weakest acid = smallest K_a
- **Amphoteric or amphiprotic** - substance can behave as an acid or a base ex. $\text{Al}(\text{OH})_3$ or H_2O
- It should be noted that Group 1 and 2 are strong bases when combined with OH^- ; Bases [OH^- combined with a metal] get weaker as you move across the periodic table from left to right

Redox and Electrochemistry

- Know the rules for determining the oxidation states
- Sum of the oxidation states in a neutral molecule must always equal **ZERO**
- **Oxidation** - loss of electrons causes the oxidation # to increase (LEO)
- **Reduction** - gaining of electrons causes the oxidation # to decrease (GER)
- To have a Redox reaction there must be a change in oxidation # and you **CANNOT** have oxidation without having reduction
- **Standard Reduction potentials show ONLY** reduction reactions, in order to change them into oxidation reactions you must reverse them and change to sign of the E° value. The strongest reducers are on the TOP of the chart and the strongest oxidizers are on the BOTTOM of the chart.
- Only metals above hydrogen will react with acids to produce Hydrogen gas
- Hydrogen is used as the standard on which the entire table is based $E^\circ = 0$ volts
- To calculate the E° of a cell first determines which one of your elements is the substance being reduced and which one is being oxidized. Use the opposite of the sign on the element being oxidized and add them up
- If E° is + then the reaction is spontaneous
- If the E° is - then the reaction is non-spontaneous
- If the $E^\circ = 0$ then the system is at equilibrium
- **Electrolytic cell** - need a battery to get going, Anode is (+) electrode & the cathode is the (-) electrode.
- **Electroplating** - plating occurs at the cathode or negative electrode. Car bumpers can be coated with protective metal in this manner. Mass increases at the site of plating.
- **Balancing Redox equations** - balance with respect to charge and mass
- The substance being reduced is the oxidizing agent

- The substance being oxidized is the reducing agent
- To figure out the E° value for a redox reaction you should use the following reaction $E^{\circ}_{\text{total}} = E^{\circ}_{\text{reduced}} + E^{\circ}_{\text{oxidized}}$

Organic Chemistry

- **Hydrocarbons** - contain only hydrogen and carbon
- **Homologous series** - successive members differ by $-\text{CH}_2$ groups.
- **Alkanes** - $\text{C}_n\text{H}_{2n+2}$ contain ALL single bonds - saturated compounds, ending is -ane
- **Alkenes** - C_nH_{2n} contain one double bond - unsaturated compounds ending are -ene
- **Alkynes** - $\text{C}_n\text{H}_{2n-2}$ contain one triple bond - unsaturated compounds ending is -yne
- **Atomic Compounds** – Based on benzene's structure (C_6H_6), toluene is related
- **Naming compounds** – follow the IUPAC rules
- As the volume of each of these homologous series increase so do their boiling points and melting points due to an increase in the van der Waals forces.
- **Properties of organic molecules** - non-electrolytes, low boiling points. & melting points, generally insoluble in polar solvents (like water), react slowly & are molecular in structure
- **Isomers** - have the same chemical formula but a different structural formula
- General formulas for functional groups:
 - Alcohols: $\text{R}-\text{OH}$
 - Organic acids: $\text{R}-\text{COOH}$
 - Ester: $\text{R}_1-\text{COO}-\text{R}_2$
 - Ketones: $\text{R}_1-\text{CO}-\text{R}_2$ [the oxygen is double bonded to a non terminal carbon]
 - Ethers: $\text{R}_1-\text{O}-\text{R}_2$ [the oxygen connects two carbon chains]
 - Aldehydes: $\text{R}-\text{COH}$ [the oxygen is double bonded to the terminal carbon]
- Organic reactions to know
 - **Addition** - adds a pair of halogens to an unsaturated hydrocarbon. One product.
 - **Substitution** - adds a halogen to a saturated hydrocarbon. Two Products.
 - **Esterfication** - acid + alcohol \rightarrow ester + water
 - **Fermentation** - $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow$ alcohol + carbon dioxide
 - **Polymerization** - $n(\text{C}_2\text{H}_4) \rightarrow (\text{C}_2\text{H}_4)_n$
 - **Combustion** - hydrocarbon + oxygen \rightarrow carbon dioxide + water
 - **Cracking** - the separation of a polymer

Nuclear Chemistry

TYPE	SYMBOL	CHARGE
Alpha Particle	${}^4_2\text{He}$	2+
Beta Particle	${}^0_{-1}\beta$	1-
Gamma Particle	γ	0

- Radioactive Nuclide Emissions
- Half-life is the time required for one half of the atoms in a radioactive nuclide to decay
- Sample decay equations
- Alpha emission
 - ${}^{210}_{84}\text{Po} \rightarrow {}^{206}_{82}\text{Pb} + {}^4_2\text{He}$
- Beta emission
 - ${}^{14}_6\text{C} \rightarrow {}^{14}_7\text{N} + {}^0_{-1}\beta$
- Gamma emission - No equation; nucleus changes from excited energy state to lower energy state
- Nuclear fission occurs when a very heavy nucleus splits into more stable nuclei of intermediate mass, releasing very large amounts of energy ${}^1_0\text{n} + {}^{235}_{92}\text{U} \rightarrow {}^{140}_{56}\text{Ba} + {}^{93}_{36}\text{Kr} + 3 {}^1_0\text{n}$
- Nuclear fusion occurs when light mass nuclei combine to form a heavier, more stable nucleus
 - $4 {}^1_1\text{H} \rightarrow {}^4_2\text{He} + 2 {}^0_{-1}\beta$
- Nuclear chemistry is used for processes such as radioactive dating; destruction of cancer cells; radioactive tracing in agriculture, medicine, and industry; and preserving food by killing bacteria and mold spores
- Nuclear waste must be isolated and allowed to decay without the radiation producing harmful effects to the environment and its inhabitants