

Chemistry Topics Reviews

(as per common core with expanded content)

I. Atomic Concepts

Atomic Theory

| | | | |
|------|---------------|---------------------------|-----------------------------------------|
| 1808 | John Dalton | Billiard ball theory | Laws (conservation, composition) |
| 1897 | JJ Thomson | Plum pudding | CRT - discovered electron |
| 1911 | E. Rutherford | Nuclear model | alpha particle / gold foil - nucleus |
| 1913 | H. Moseley | Protons - Z | x-ray absorption edges |
| 1913 | Niels Bohr | Quantum Mechanics | electrons shells / spectra |
| 1924 | L. de Broglie | e^- wave theory | explained energy levels (e^- shells) |
| 1925 | W. Pauli | Pauli exclusion principle | each e^- has unique quantum numbers |
| 1926 | Schrödinger | Orbitals | probability function for electrons |
| 1932 | J. Chadwick | Neutrons | Be + alpha particle |

Atom

| Particle | Mass | Charge |
|--------------------|----------|--------|
| proton (p^+) | 1 u | +1 |
| neutron (n^0) | 1 u | 0 |
| electron (e^-) | 1/1830 u | -1 |

Nucleus

protons (p^+)

neutrons (n^0)

Electrons (e^-)

Enclose a lot of empty space surrounding the nucleus

Found in orbitals (region of most probable location) as calculated by Schrödinger

Have a distinct amount of energy (Pauli exclusion principle)

Ground state: all e^- are in the lowest possible energy states (e^- shells)

Excited state: at least one e^- has gained energy and moved to a higher e^- shell

Electrons moving from a higher energy level (e^- shell) to a lower one (closer to the nucleus) emit a specific amount of energy (color of light) – spectra as per N. Bohr

Valence electrons: e^- in the outermost shell (energy level)

The number of valence e^- defines chemical behavior (properties)

Can be read off the Periodic Table as the electron configuration (energy level diagram)

Example: Ca = 2–8–8–2 therefore Ca has 2 valence e^- (found in Group II A)

Isotopes: atoms of the same element (having the same number of p^+) but differing numbers of neutrons (differing mass numbers)

Atomic number: the number of protons in the nucleus of an atom (Z)

Mass number: the number of protons and neutrons in a specific isotope (nuclide)

Atomic mass: the weighted average of the masses of the naturally occurring isotopes

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II. Periodic Table

Elements on the Periodic Table are arranged in order of increasing atomic number

Atomic number (Z): identifies an element

Mass number: identifies a particular isotope (nuclide)

Common notations representing isotopes include:

${}^{14}_6\text{C}$, ${}^{14}\text{C}$, carbon-14, C-14

The placement of an element on the Periodic Table (period and group) indicates the chemical and physical properties of that element

Elements can be classified as

Metals: about 85% of the elements on the Periodic Table (found on the left side)

Metalloids: (B, Si, Ge, As, Sb, and Te)

Nonmetals: elements up and to the right of the metalloids

Noble gases: Group 18 (or Group VII A)

Elements can be differentiated by their physical properties (density, conductivity, malleability, solubility, and hardness)

Elements can be differentiated by their chemical properties (properties that describe how an element behaves during a chemical reaction)

Some elements exist in two or more forms in the same phase (allotropes)

allotropes differ in their molecular structures (crystal structures) and therefore in properties

Groups 1, 2, and 13-18 (Groups IA, IIA, IIIA, IVA, VA, VIA, VIIA, VIIIA)

elements in the same group (except He) have the same number of valence electrons

therefore elements in the same group have similar chemical properties

The succession of elements within the same group (family) demonstrates characteristic trends

examples: atomic radius, ionic radius, electronegativity, first ionization energy,

metallic/nonmetallic properties

The succession of elements across the same period (series) demonstrates characteristic trends

of atomic radius, ionic radius, electronegativity, first ionization energy,

metallic/nonmetallic properties

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III. Moles / Stoichiometry

Compound: a substance composed of two or more elements combined in a fixed proportion

Compounds can be broken down by chemical means

Elements cannot be broken down by chemical means

Because compounds have a fixed proportion, they can be represented by a chemical formula

IUPAC assigns a specific name to a given formula

Types of formulas

Empirical: simplest whole number ratio

Molecular: actual ratios of atoms in a molecule

Structural: a representation that shows all bonding in a molecule

Chemical reactions are bound by conservation of

Mass

Energy

Charge

Balanced chemical equations represent the conservation of atoms

Coefficients in a balanced equation can be used to determine mole ratios in the reaction

Formula mass: the sum of the atomic masses of atoms in a formula

Molar mass (or gram formula mass) represents one mole of that substance

Percent composition by mass of each element in a formula can be calculated

Types or classes of reactions include:

Synthesis $A + X \rightarrow AX$

Decomposition $AX \rightarrow A + X$

Single replacement $A + BX \rightarrow AX + B$

Double replacement $AX + BY \rightarrow AY + BX$

Combustion $C_xH_{2y} + (x+1/2y) O_2 \rightarrow y H_2O + x CO_2$

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IV. Chemical Bonding

Compounds can be differentiated (separated or identified) by their chemical and physical properties

One method to classify compounds is to place them in two major categories

Ionic compounds: usually formed by reacting a metal and a nonmetal

Compounds with polyatomic ions have both ionic and covalent bonds

Covalent compounds: usually formed by reacting two nonmetals

Chemical bonds are formed when valence electrons are:

Mobile within metals, sea of electrons (metallic bonds)

Transferred from one atom to another (ionic bonds)

Ions form when atoms gain or lose electrons

Cations: positive ions formed when atoms lose electrons will be smaller than the atom

Anions: negative ions formed when atoms gain electrons will be larger than the atom

Shared between atoms forming molecules (covalent bonds)

Molecular polarity is determined by molecular shape and charge distribution

Symmetrical (nonpolar) molecules include: CO_2 , CH_4 , diatomic elements (Br_2 or N_2)

Asymmetrical (polar) molecules include: HCl , NH_3 , and H_2O

Chemical bonds are stored (or potential) energy

Bond breaking is always endothermic (energy is absorbed and appears as a reactant)

Bond formations is always exothermic (energy is released and appears as a product)

Atoms attain a stable valence electron configuration by bonding with other atoms

Usually a stable octet is formed (exceptions include H, for example)

Noble gases already have a stable valence configuration and tend not to bond

Physical properties of substances can be explained in terms of chemical bonds (metallic, ionic, or covalent) and intermolecular forces

These properties include: conductivity, malleability, solubility, hardness, melting point, boiling point, and conductivity (electrical and thermal)

Intermolecular forces are created by unequal charge distribution result in varying degrees of attraction between molecules

Hydrogen bonds strongest intermolecular bond requiring N, O, or F atoms

Dipole–dipole interactions permanent polarization (determined by shape and polarity)

van der Waals forces temporary polarization caused by electron movement

Types of solids:

Metallic: characterized by a sea of mobile electrons

Ionic: characterized by electron transfer forming ions (+ and –)

Network: characterized by a 3-D network of covalent bonds

Molecular: characterized by their intermolecular forces

Lewis electron-dot diagrams (or structures) can represent valence electron arrangements in elements, compounds, and ions

Electronegativity indicates how strongly bonded atoms attract electrons in a chemical bond with electronegativity values assigned arbitrarily (Linus Pauling)

Electronegativity differences is used to assess the degree of bond polarity

$\Delta\text{EN} > 1.7$ Ionic bond

$1.7 \geq \Delta\text{EN} \geq 0.4$ Polar covalent bond

$\Delta\text{EN} < 0.4$ Nonpolar covalent bond

Electrons in nonpolar molecules can be polarizable and form van der Waals forces

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V. Physical Behavior of Matter

Matter is classified as:

Pure substances have constant composition throughout a sample and from sample to sample

Elements: composed of atoms with the same atomic number throughout

Cannot be broken down by chemical means

Compounds: composed of two or more elements in constant proportions

Can be broken down by chemical means

Mixtures: (heterogeneous or homogeneous) have varying compositions and each component retains its original properties

Can be separated by physical means based on such properties as density, particle size, molecular polarity, boiling point, freezing point, and solubility

Solution: homogeneous mixture of a solute and a solvent

Solubility of a solute depends upon temperature, pressure, and the chemical nature of both the solute and the solvent

Concentration can be expressed as molarity (M), percent by volume, percent by mass, or parts per million (ppm)

Phases of matter:

Solids

Liquids

Gases

The structure and arrangement of particles and their interactions determine the physical state of a substance at a given temperature and pressure

The addition of a nonvolatile solute to a solvent causes:

Boiling point elevation

Freezing point depression

Such colligative properties depend solely on the concentration of solute particles

Electrolytes (ionic compounds) contribute multiple particles (*i*, van't Hoff factor)

Energy has multiple forms:

Chemical, electrical, electromagnetic (light), thermal, mechanical, and nuclear

Heat is the transfer of energy (usually thermal energy which is associated with the random motions of atoms and molecules) from a body of higher temperature to a body of lower temperature

Temperature: a measure of the average kinetic energy of the particles in a sample

Temperature is not a form of energy (it is a measure of a particular form of energy)

Ideal gas: a model used to explain the behavior of real gases

Gases are most ideal at high temperatures and low pressures

Kinetic Molecular Theory (KMT) of an ideal gas:

1. Volume of gas particles is negligible compared to the distance separating them
2. Particles are in constant, rapid, random, straight-line motion
3. Undergo perfectly elastic collisions (energy transfers result in no loss of total energy)
4. No attractive (or repulsive) forces act between particles
5. $KE = \frac{1}{2} mu^2$

KMT describes the relationships of pressure, volume, temperature, velocity, and frequency and force of collisions among gas molecules

Combined gas law: $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

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Equal volumes of different gases at the same temperature and pressure contain an equal number of particles (Avogadro's hypothesis)

The concepts of kinetic and potential energy can be used to explain physical processes that include: fusion (melting), vaporization (boiling, evaporation), condensation, sublimation, and deposition

Both chemical and physical changes can be endothermic or exothermic

Physical changes result in rearrangement of existing particles of existing substances

Chemical changes result in the formation of different substances (change chemical identity)

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VI. Kinetics / Equilibrium

Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation

Rate of a chemical reaction depends upon several factors:

Temperature: measure of the average kinetic energy, doubles rate every 10K

Concentration: usually in molarity (M) and most used for gases and solutions

Nature of the reactants: in terms of chemical activity

Surface area: usually for solids

Presence of a catalyst or inhibitor: these change E_a by changing the reaction pathway

Some chemical and physical changes can reach equilibrium (always a dynamic state)

Equilibrium occurs when the rate of the forward reaction equals the rate of the reverse reaction

The measurable quantities of reactants and product remain constant at equilibrium

Equilibrium requires a closed system

Le Châtelier's principle predicts the effects of stress or changes in pressure (only when the number of moles of gases varies between reactants and products), volume, concentration, and temperature of a system at equilibrium

Energy released or absorbed by a chemical reaction can be represented by a potential energy diagram

Heat of reaction (energy released or absorbed during a chemical reaction) is equal to the difference between the potential energy of the products and the potential energy of the reactants

Catalysis provides an alternate reaction pathway (or mechanism) which lowers the activation energy (E_a) of a reaction

Entropy is a measure of the number of degrees of freedom (randomness or disorder) of the particles in a system

Systems in nature tend to undergo changes toward lower energy and higher entropy

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VII. Organic Chemistry

Organic compounds contain carbon atoms

Carbonates, carbides, and carbon oxides are not considered organic molecules

Carbon can:

Form chains (catenate)

Form rings

Form networks

Form a variety of other structures

Organic compounds are named using the IUPAC system

Hydrocarbons contain only H and C atoms

Saturated hydrocarbons contain only carbon-carbon (C-C) single bonds

Unsaturated hydrocarbons contain at least one multiple (double or triple) C-C bond

In a multiple covalent bond, more than one pair of electrons are shared between two atoms

Homologous series

| | | | |
|---------|---------------------------|-----------------|-------------|
| Alkanes | single C-C bonds only | $C_nH_{(2n+2)}$ | saturated |
| Alkenes | contain a double C=C | $C_nH_{(2n)}$ | unsaturated |
| Alkynes | contain a triple C≡C bond | $C_nH_{(2n-2)}$ | unsaturated |

Functional groups: impart distinctive physical and chemical properties to organic compounds

| | | | | |
|---------------|----------|--------|-----------|---------|
| halides | alcohols | ethers | aldehydes | ketones |
| organic acids | esters | amines | amides | |

Isomers: organic compounds having the same molecular formulas but different structures and properties

Types of organic reactions:

| | | | | |
|--------------|----------------|--------------|----------------|----------------|
| addition | elimination | substitution | polymerization | esterification |
| fermentation | saponification | combustion | | |

VIII. Oxidation-Reduction

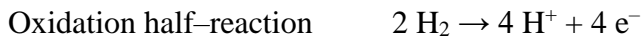
Redox reactions (oxidation-reduction reactions) involve the transfer of electrons

Reduction is the gain of electrons (the charge is reduced)

Oxidation is the loss of electrons

The number of electrons lost must equal the number of electrons gained (conservation of charge)

Oxidation-reduction reactions can be written as half-reactions



Oxidation numbers (states) can be assigned to atoms and ions

Changes in oxidation numbers indicate that oxidation and reduction have occurred

Electrochemical cells can be voltaic (spontaneous) or electrolytic (nonspontaneous)

Voltaic cells spontaneously convert chemical energy to electrical energy

Electrolytic cells require energy to produce a chemical change that stores potential energy

In all electrochemical cells:

Oxidation occurs at the anode

Reduction occurs at the cathode

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IX. Acids, Bases, and Salts

Behavior of many acids and bases can be explained by the Arrhenius theory

Arrhenius acids and bases are electrolytes (a substance which when dissolved in water forms a solution capable of conducting electricity)

The ability to conduct electrical current in solution depends on the concentration of ions

Arrhenius acid: produces $\text{H}^+_{(\text{aq})}$ as the only positive ion in an aqueous solution

$\text{H}^+_{(\text{aq})}$ can also be written as $\text{H}_3\text{O}^+_{(\text{aq})}$, the hydronium ion

Arrhenius base: produces $\text{OH}^-_{(\text{aq})}$ as the only negative ion in an aqueous solution

There are other acid base theories:

Brønsted-Lowry

Acid: a proton (H^+) donor

Base: a proton (H^+) acceptor

Lewis

Acid: an electron pair acceptor

Base: an electron pair donor

Acidity or alkalinity of a solution can be measured using the pH scale or by using indicators

The pH scale is exponential

$$\text{pH} = -\log[\text{H}^+_{(\text{aq})}]$$

An increase of one unit (say 3 to 4) on the pH scale represents a tenfold decrease in $[\text{H}^+_{(\text{aq})}]$

In the process of neutralization, an Arrhenius acid and an Arrhenius base react to form a salt and water

Titration: laboratory process using a buret in which the volume of a solution of known concentration is used to determine the concentration of another solution

Endpoint: the condition in a titration where the $[\text{H}^+_{(\text{aq})}] = [\text{OH}^-_{(\text{aq})}]$

Titration of a weak acid and a strong base will give a basic endpoint

(See Table M on the 2011 Reference Tables for Chemistry to find a basic indicator)

Titration of a strong acid and a weak base will give an acidic endpoint

Titration of a strong acid and a strong base will give a neutral endpoint

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X. Nuclear Chemistry

Nuclear reactions include natural transmutation, artificial transmutation, fission, and fusion

Nuclear reactions occur because some isotopes have unstable nuclei (radioisotopes)

Stability of an isotope is based on the ratio of neutrons and protons in the nucleus

Light elements $Z < 20$ $p^+ : n^0 = 1 : 1$

Heavy elements $Z < 70$ $p^+ : n^0 = 1 : 1.5$

Transmutation: a change in the nucleus of an atom that converts it from one element to another element (with a different atomic number, Z)

Natural transmutation: spontaneous nuclear decay

Artificial transmutation: nuclear decay induced by high-energy particle bombardment

Unstable isotopes spontaneously decay emitting radiation

| | | | | |
|--------------------------|--------------------------------------------|--------|-----------------|-----------------------|
| Alpha (α) decay | ${}^4_2\text{He}$ | slow | +2 charge | low penetrating power |
| Beta (β) decay | ${}_{-1}^0\text{e}$ or ${}_{+1}^0\text{e}$ | medium | +1 or -1 charge | medium penetration |
| Gamma (γ) decay | ${}^0_0\gamma$ | fast | 0 charge | high penetration |

Each radioisotope has a specific mode and rate of decay (half-life)

Energy released during nuclear reactions comes from the fractional amount of mass converted into energy by Einstein's famous equation: $E = mc^2$

Mass defect: the small amount of mass missing from nucleons (protons and neutrons) of heavier elements due to the conversion of mass to energy to bind protons together in the nucleus (in order to overcome the electrostatic repulsion of the protons)

Binding energy: the energy equivalent of the mass defect (given by $\Delta E = \Delta mc^2$)

There are risks and benefits associated with fission and fusion reactions

Beneficial uses: radioactive dating (${}^{14}\text{C}$ and ${}^{238}\text{U}$), tracing chemical and biological processes, industrial measurement, nuclear power, detection (${}^{131}\text{I}$) and treatment of disease (${}^{60}\text{Co}$)

For nuclear power, the energy released during nuclear reactions is much, much greater than the energy released during chemical reactions

Risks: biological exposure, long-term storage and disposal, and nuclear accidents

Nuclear reactions can be represented by equations that include symbols which represent nuclei (including mass number and atomic number: ${}^{60}_{27}\text{Co}$), subatomic particles (with mass number and charge, see alpha and beta particles above), and/or emissions such as gamma radiation (${}^0_0\gamma$)