

Chapters 1: Introductions & The Chemistry Of Measurement

I. Introduction

- A. Definition Of Chemistry
- B. Organization Of Chemistry
- C. Study Of Chemistry (The Scientific Method)

II. The Science Of Measurement

- A. Importance Of Measurement
- B. Conditions For Meaningful Measurement
 - 1. Define
 - 2. Standardize
 - 3. Compare
- C. Content Of Measurements
 - 1. Numbers (Uncertainty, Sig. Figs., Exponents & Equations)
 - 2. Units (Systems, Standards & Derived Units)
- D. Equivalent Relationships
 - 1. Dimensional Analysis
 - 2. Percentage As A Conversion Factor
 - 3. Temperature Conversions
 - 4. Energy Units And Conversions

Significant Figures

- All Nonzero Digits Are Sig.. Figs. 28475 5 Sig.. Figs.
- All Enclosed Zeros Are Sig. Figs. 300,452 6 Sig.. Figs.
- Leading Zeros Are Not Sig. Figs. 0.00082 2 Sig.. Figs.
- Trailing Zeros With Decimal Are Sig. Figs. 34,000.00 7 Sig. Figs.
- Trailing Zeros Without Decimal Are Not Sig. Figs. 96,000,000 2 Sig. Figs.
- Sig. Figs. In Multiplication And Division: The Number Of Sig. Figs. In The Final Answer Should Equal The Number Of Sig.Figs Of The Given Data Number Having The Least Number Of Sig. Figs.
- Sig. Figs. In Addition And Subtraction: Convert All Data To Same Order Of Magnitude And Add Or Subtract Relative The The Decimal Point. Express Answer Relative To The Given Data Having The Least No Of Sig. Figs.

Chapter 2: Atoms, Molecules & Ions **(The Organization Of Matter)**

I. Description Of Matter

- A. Classification Of Matter (Mixtures & Substances)**
- B. Study of Matter via Properties Of Matter (System & Surroundings)**

II. Atomic Structure

- A. Fundamental Particles**
- B. Atomic Structure From The Periodic Chart**
- C. Variations In Atomic Structure**
 - 1. Isotopes (Same At. No. But Different At. Mass)**
 - 2. Ions (Charged Particles Due To Gain Or Loss Of Electrons)**

III. The Periodic Table

- A. Periods & Groups**
- B. Metals, Nonmetals & Metalloids**
- C. Block Groupings Of Elements (Representative, Transition & Inner Transition Elements)**

IV. Atomic Theory & Laws Of Chemical Composition

- A. Law Of Definite Proportions**
- B. Elemental Percent Composition**

V. Chemical Formulas & Nomenclature

- A. Formula Classification (Ionic & Molecular Compounds)**
- B. Types Of Formulas (Empirical, Molecular & Structural)**
- C. Writing Formulas**
- D. Nomenclature**

Chapter 3: Chemical Reactions

I. The Chemical Equation

- A. The Concept Of Mass Balance
- B. Classification Of Chemical Reactions
 - 1. Combination Reactions
 - 2. Decomposition Reactions
 - 3. Single Replacement Reactions
 - 4. Double Replacement Reactions (= Metathesis Rxns)

ii. Ions In Aqueous Solution

- A. Electrolytes
 - 1. Type Of Electrolytes (*Strong, Weak & Nonelectrolytes*)
 - 2. Acids And Bases
 - A. Theories Of Acids And Bases
 - B. Strong Vs`Weak Acids
 - C. Molecular And Ionic Equations

iii. Reactions In Aqueous Solution

- A. The Driving Force Of Chemical Reaction
- B. Precipitation Reactons
- C. Neutralization Reactions
- D. Reactions Forming Gas Phase Products

DECOMPOSITION REACTIONS

CARBONATES, METALLIC \longrightarrow METAL OXIDES + NON-METAL OXIDES

OXIDES OF METALS \longrightarrow METAL + OXYGEN

ACIDS, WEAK OXO-ACIDS \longrightarrow NON-METAL OXIDE + WATER

CHLORATES, METALLIC \longrightarrow SALT + OXYGEN

HYDROXIDES, METALLIC \longrightarrow METALLIC OXIDES + WATER

ELECTROLYTIC PRODUCTS \longrightarrow BASIC ELEMENTS IN STANDARD STATES

Driving Force Of Metathesis Reactions

Solubility Of Ionic Compounds:

<u>Soluble</u>	<u>Except</u> These Salts Are Insoluble			
Halides	Hg ₂ ⁺²	Ag ⁺	Pb ⁺²	
Ammonium Salts	All Soluble			
Acetate Salts	All Soluble			
Nitrate Salts	All Soluble			
Group Ia	All Soluble			
Sulfates	Ba ⁺²	Ca ⁺²	Sr ⁺²	Pb ⁺²

<u>Insoluble</u>	<u>Except</u> These Salts Are Soluble			
Hydroxides	GpIA	NH ₄ ⁺	Ba ⁺²	Sr ⁺²
Phosphates	GpIA	NH ₄ ⁺		
Carbonates	GpIA	NH ₄ ⁺		
Sulfides	GpIA	NH ₄ ⁺	GpIIA	

Strong & Weak Acids:

Strong Acids: HCl, HBr, HI, HNO₃, HClO₄, H₂SO₄

Weak Acids: Any Compound Containing Acidic Hydrogens Not A Member Of The Above Set Of Acids.

Gas Decomposition Products:

Gas Product	Metathesis Reaction	Decomposition Reaction
CO ₂	CaCO ₃ + 2HCl → CaCl ₂ + H ₂ CO ₃	H ₂ CO ₃ → CO ₂ + H ₂ O
NO & NO ₂	NaNO ₂ + HCl → NaCl + HNO ₂	2HNO ₂ → NO + NO ₂ + H ₂ O
SO ₂	Na ₂ SO ₃ + 2HCl → 2NaCl + H ₂ SO ₃	H ₂ SO ₃ → SO ₂ + H ₂ O
H ₂ S	Na ₂ S + 2HCl → H ₂ S + 2NaCl	None / H ₂ S Is Gas Product
NH ₃	NH ₄ Cl + NaOH → NaCl + NH ₄ OH	NH ₄ OH → NH ₃ + H ₂ O

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Chapter 4: Chemical Stoichiometry (Formulas, Equations & Solutions)

I. Mass And Moles

- A. Molecular Weight And Formula Weight**
- B. The Mole Concept**

ii. Determination Of Chemical Formula

- A. Mass Percentages From Formulas**
- B. Elemental Analysis**
- C. Determining Empirical & Molecular Formulas**

iii. Stoichiometry & Quantitative Relations In

Chemical Equations

- A. Mole Interpretation Of An Equation**
- B. Stoichiometry Of A Chemical Reaction**
- C. Limiting Reactant Problems**
- D. Theoretical & %Yields**

iv. Calculations Involving Solutions

- A. Expressing Concentration Of Solution**
- B. Dilution Process & The Dilution Equation**
- C. Stoichiometry Of Solution Reactions**

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Chapter 5: The Gaseous State

- I. Stoichiometry Of Gas Phase Reactions**
 - A. Avogadro's Law & Molar Volume**
 - B. Reactive Volumes**

- ii. The Empirical Gas Laws**
 - A. Boyles Law**
 - B. Charles Law**
 - C. Gay Lussac Law**
 - D. Combined Gas Law**
 - E. Ideal Gas Law**

- lii. Gas Pressure And Its Measurement**
 - A. Concept Of Gas Pressure**
 - B. Measurement Of Gas Pressure**
 - C. The Gas Law Variables**

- iv. Dalton's Law Of Partial Pressures**

- V. Graham's Law Of Effusion**

- Vi. Kinetic-Molecular Theory Of Gasses**

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Chapter 6: Thermochemistry

- I. Concept of Energy and Energy Change**
 - a. Definition**
 - b. Classification (Practical vs Physical)**
 - c. Law of Conservation of Energy**
 - d. Energy Units**

- II. Heat of Reaction (Change in Internal Energy)**
 - a. Positional Energy (Energy Content) and Energy Change**
 - b. Energy Flow**
 - i. System and Surroundings**
 - ii. Endothermic and Exothermic Process**
 - iii. Identifying Endothermic and Exothermic Process**

- III. Thermochemical Process**
 - a. Forms and Notation of Thermochemical Equations**
 - b. Stoichiometry of Thermochemical Equations**
 - i. Proportional Relationships**
 - ii. Standard Heats of Formation**
 - iii. Hess's Law**

Chapter 11: States of Matter

OBJECTIVE: EXPLAIN THE NATURE AND BEHAVIOR OF SOLIDS, LIQUIDS & GASSES IN TERMS OF THE KINETIC-MOLECULAR THEORY

INTRODUCTION:

Comparison Of Solids, Liquids & Gases
The Kinetic-Molecular Theory

CHANGE OF STATE:

Phase Changes:

The Transition Triangle (Terminology)

Vapor Pressure

Boiling Point & Melting Point

Heating Curve

Clausis-Clapeyron Equation

The Phase Diagram

Melting Pt Curve, Boiling Pt Curve & Sublimation Curve

Triple Point

Critical Point

INTERMOLECULAR FORCES:

Dipole-Dipole Forces, H-Bonding Forces & Van Der Waal's Or London Forces

THE LIQUID STATE: PROPERTIES OF LIQUIDS (SURFACE TENSION & VISCOSITY)

SURFACE TENSION = The work force needed to increase surface area.

FACTORS AFFECTING SURFACE TENSION = Soaps, Detergents and Surfactants

CAPILLARY ACTION = Adhesion & Cohesion

VISCOSITY = Fluid resistance to flow (Fluid Thickness)

Viscosity decreases with increasing temperature

Motor Oil = SAE 10W/40 Society of Automotive Engineers (Viscosity of 10 in Winter and 40 in Summer)

THE SOLID STATE:

CLASSIFICATION OF SOLIDS

MOLECULAR SOLIDS: A solid that consists of atoms and molecules held together by intermolecular physical forces. Examples include ice, solid CO_2 (dry ice), etc...

METALLIC SOLIDS: A solid that consists of positive cores of atoms held together by a "sea of electrons" (i.e., metallic bonding). Examples include iron, copper, silver, aluminum etc...

IONIC SOLIDS: A solid consisting of cations and anions held together by electrostatic interaction of opposite charges; i.e., ionic bonding. Examples include ionic salts composed of metal cations and nonmetal or polyionic anions such as NaCl , CaCl_2 , CoSO_4 , etc...

COVALENT NETWORK SOLIDS: A solid that consists of atoms held together in large network structures or chains by covalent bonds. Examples include diamond, graphite and asbestos.

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Chapter 12: Solutions

I. Types of Solutions

- A. Definition of Solutions**
- B. Components of A Solution**
- C. Types of Solutions**
- D. Character of Solutions**

ii. Terminology of Solutions

iii. Solubility and The Solution Process

- A. Factors Affecting Solubility**
- B. Molecular Solutions**
- C. Ionic Solutions**

iv. Colligative Properties

- A. Concentration Expressions**
- B. Vapor Pressure Of Solutions**
- C. Boiling - Point Elevation and Freezing - Point Depression**
- D. Osmosis**
- E. Colligative Properties of Ionic Solutions**

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Chapter 12: Solutions (Continued)

Terminology Of Solutions

Solubility

Factors Affecting Solubility
(Tendency Of Solute & Solvent To Mix)
(Tendency Toward Lowest Energy State)
(Structure, Temperature, Pressure)

Units Of Solubility

Saturated Solutions

Equilibrium

Unsaturated Solutions

Dilute Solutions

Concentrated Solutions

Miscibility
(Miscible, Partially Miscible, Immiscible)

Aqueous Solutions

Chapter 7 - 8: Atomic Theory

I. Introduction

- A. Studying The Structure Of Matter
- B. History Of The Modern Atomic Theory

ii. Basic Atomic Structure

- A. Subatomic Particles
- B. Structure Of Atoms
- C. Variations In Atomic Structure
- D. Atomic Dimensions

iii. Theories Of Atomic Structure

- A. The Rutherford Model (The Shell Model)
- B. The Bohr Model (The Concentric Ring Model)
- C. The Schrodinger Model (The Modern Quantum Model)

Iv. Electron Configurations

- Linear Atomic Dimensions:

Given The Diameter Of The Nucleus Of An Atom To Be 1.00 Cm.

(Approximate Size Of A Small Marble), What Is The Radius Of The Atom Relative To A Nuclear Diameter Of This Size?

(Note: Avg. Atomic Diameter = 10^{-8} cm. And Avg Nuclear Diameter= 10^{-13} cm.)

- Atomic Mass Relationships:

The Absolute Atomic Mass Of Helium Is 6.6978×10^{-24} Gm. Calculate The Weight Percentage Represented By The Nucleus Weighing 6.696×10^{-24} Gm. And The Electron Cloud (Consisting Of 2 Electrons) Weighing 1.82×10^{-27} Gm.

- Atomic Volume Relationship:

How Many Atomic Nuclei Would Be Required To Fill A Spherical Atomic Volume Of 5.233×10^{-25} Cm.³ Assuming The Nucleus Is Spherical And Has A Diameter Of 10^{-13} Cm. ?

(Note: Volume Of Sphere = $\frac{4}{3}\pi R^3$)

Chapter 7 - 8: Atomic Theory (Continued)

Historical Discovery Of Subatomic Particles

I. The Electron

A. J.J. Thompson (1897) - Cathode Ray Tube:Electron Has Mass And Charge.

B. Millikan (1909) - Oil Drop Experiment: Determined The e/m Ratio For The Electron = $1/1837$ Of Hydrogen Atom Mass.

ii. The Proton

A. Goldstein (1886) - Cathode Ray Tube With Perforated Cathode. Measured Relative Q/M (Charge To Mass Ratio) of Protons And Electrons. (Mass $P^+ = 1.0073$ Amu)

B. Becquerel (1896) - Discovered Radioactivity or, The Spontaneous Decomposition Of Atomic Nuclei.

C. Rutherford (1919) - High Velocity α -Particles Bombarded Nitrogen & Aluminum To Find Proton Emissions.

iii. The Neutron

A. James Chadwick (1932) - Bombarded Beryllium With α -Particles. Discovered Emission Of "Uncharged" Particle. Named Particles "Neutrons". (Mass = 1.0087 Amu).

POSTULATES FOR ATOMIC MODELS OF THE ATOM

Rutherford Model Of The Atom

Experimental Postulates:

1. A Very Dense, Small Nucleus Exists In The Center Of The Atom. This Nucleus Contains Most Of The Mass Of The Atom And All Of The Positive Charge.
2. Electrons Occupy Most Of The Total Volume Of The Atom And Are Located Outside The Nucleus.
3. When An Alpha Particle Scores A Direct Hit On A Nucleus, It Is Deflected Back Along The Incoming Path.
4. A Near Miss Of A Nucleus By An Alpha Particle Results In Repulsion And Deflection.
5. Most Of The Alpha Particles Pass Through Without Any Interference, Because Most Of The Atomic Volume Is Empty Space.
6. Electrons Have So Little Mass That They Do Not Deflect The Much Larger Alpha Particles.

The Bohr Model Of The Atom

Experimental Postulates:

1. The Electron In A Hydrogen Atom Travels Around The Nucleus In A Circular Orbit.
2. The Energy Of The Electron In An Orbit Is Proportional To Its Distance From The Nucleus. The Further The Electron Is From The Nucleus, The More Energy It Has.
3. Only A Limited Number Of Orbits With Certain Energies Are Allowed; Ie, The Orbits Are Quantized At Discrete Energy Values.
4. The Only Orbits That Are Allowed Are Those For Which The Angular Momentum Of The Electron Is An Integral Multiple Of Planck's Constant Divided By 2π .
5. Light Is Absorbed When An Electron Jumps To A Higher Energy Orbit And Emitted When An Electron Falls Into A Lower Energy Orbit.
6. The Energy Of The Light Emitted Or Absorbed Is Exactly Equal To The Difference Between The Energies Of The Orbits.

Schrodinger Model Of The Atom:

Experimental Postulates

1. Electrons Have A Wave - Particle Duality; That Is, They Demonstrate Both Wave And Particle Characteristics As Defined By The De Broglie Relation ($\lambda = h/mc$).
2. Both Position And Momentum Of The Electron About The Nucleus Can Not Be Defined With Certainty As Defined By Heisenberg's Uncertainty Principle [$(\Delta X)(\Delta P) \geq h / 4\pi$].
3. The Position Of The Electron About The Nucleus Is Defined In Terms Of A Statistical Probability Of Finding The Electron Within A Specified Region As Described By Probability Wave Functions (ψ)².
4. The Behavior Of The Electron About The Nucleus Of The Atom Is Described By A Set Of Four Quantum Numbers That Define The Energy State Of Each Electron. The Pauli Exclusion Principle States That No Two Electrons Have The Same Set Of 4-Quantum Numbers.
5. The Quantum Numbers Are:
 - Principle (N) = Energy Level (Location Of Electron)
 - Orbital (L) = Shape Of Suborbital (S,P,D,F)
 - Magnetic (M) = Orientation Of Suborbital
 - Spin (M_s) = Rotation Of Suborbital

QUANTUM NUMBERS:

NAME	SYMBOL	DISCRIBES	VALUES
Principle	n	Location / Energy Level	n = 1,2,3,4,5,6,7
Orbital	l	Shape of Sub-levels	l = 0,1,2,3 ~ s,p,d,f
Magnetic	m	Orientation of Sub-levels	m = (see below)
Spin	m _s	Rotation	m _s = + 1/2 , - 1/2

Sub-level	-3	-2	-1	0	+1	+2	+3
s				s			
p			p _x	p _y	p _z		
d		d ₁	d ₂	d ₃	d ₄	d ₅	
f	f ₁	f ₂	f ₃	f ₄	f ₅	f ₆	f ₇

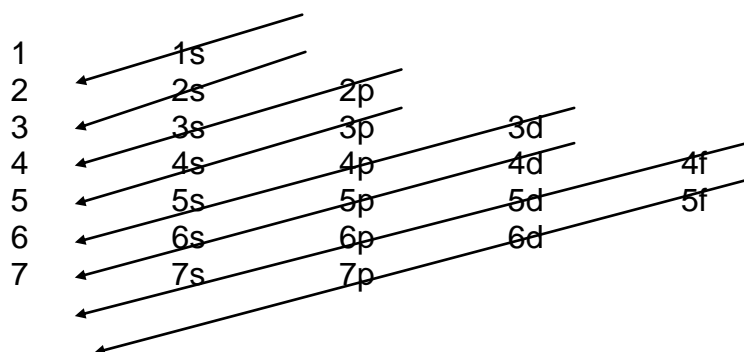
RULES FOR ASSIGNING QUANTUM NUMBERS:

- Each electron has 4 quantum numbers.
- No two electrons in any one atom may have the same set of 4 quantum numbers.
- The maximum number of electrons per energy level is defined by $2n^2$.
- The maximum number of electrons per sub-level is (2 x #orientations):
 - s = 2 e^{-s} in 1-orientation
 - p = 6 e^{-s} in 3-orientations
 - d = 10 e^{-s} in 5-orientations
 - f = 14 e^{-s} in 7-orientations
- The maximum number of electrons per sub-level orientation is = 2.

THE AUFBAU PRINCIPLE:

Electrons enter sub-level orientations having the lowest available energy.

THE AUFBAU DIAGRAM:



Chapter 9 - 10: Chemical Bonding And Molecular Geometry

I. Introduction To Bonding

- A. Definition & Classification Of Chemical Bonds
- B. Atomic Structure & Bonding
(Valence, Ionization, Oxidation, & Ionic Radius)
- C. Describing Chemical Bonds
(Dot Structures, Octet Rule, Electron Configurations & Periodic Trends)

II. Ionic & Covalent Bonding

- A. Ionic Bonding
 - 1. Formation And Structure (Ion Pairs)
 - 2. Factors Affection Ionic Bonding
 - Low Ionization Energy For Cation.
 - High Electron Affinity For Anion
 - High Lattice Energy
 - 3. Energy Transitions In Ionic Bonding (The Borne-Haber Cycle)
- B. Covalent Bonding
 - 1. Formation And Structure (Electron Pairs)
 - 2. Energy Factors (Bond Energy And Bond Length)
 - 3. Writing Electron Dot Structures (AVOBEC Method)
 - 4. Exceptions To The Octet Rule
 - 5. Coordinate Covalent Bonds

III. Dynamic Factors in Bonding & Molecular Structures

- A. Electronegativity
- B. Bond Polarity
- C. Molecular Polarity
- D. Formal Charge
- E. Resonance

IV. Molecular Geometry

- A. Molecular Shapes
(Linear, Trigonal Planar, Tetrahedral, Trigonal Bipyramid, Octahedral, Pentagonal Bipyrimid)
- B. Theories Of Molecular Shapes
 - 1. VSEPR Theory
 - 2. VB-Theory
 - 3. MO-Theory

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Course Contents

The Scientific Method

Problem/Data/Hypothesis/Testing/Conclusions

The Science Of Measurement

Content Of Measurements

Systems Of Measurement

Dimensional Analysis

The Organization Of Chemistry

Mixtures, Compounds & Elements

Introduction To The Structure Of Matter

Basic Atomic Structure

The Periodic Table

Naming Chemical Compounds

Chemical Reactions And Stoichiometry

Chemical Rxns And Concept Of Mass Balance

Chemical Stoichiometry (The Mole Concept)

Gas Laws And Thermochemistry

Empirical Gas Laws And The Ideal Gas Law

Chemical Energy From Reactions

States Of Matter

Solids, Liquids & Gases

Solutions

Historical Development Of The Atomic Theory - Models Of The Atom

Electronic Structure Of The Atom

Chemical Bonding

Ionic Bonding

Molecular Bonding

Molecular Geometry

CONVERSION FACTORS:

LENGTH	WEIGHT	TIME	VOLUME	ENERGY
1 mi. = 5,280ft. 1 mi. = 1,760yds. 1 mi. = 1,610 m. 1 yd. = 3.0 ft. = 91.4cm. 1 ft. = 12 in. 1 in. = 2.54 cm. 1 m. = 100 cm. 1 Km. = 1000 m. = 0.62 mi. 1 m. = 39.37 in.	1 Ton (T) = 2000 lbs. 1 lb. = 454 gms. 1 lb. = 16 oz. 1 Kg. = 1000 g. 1 oz. = 28.3 g. 1 Kg. = 2.205 lb.	1 min. = 60 sec. 1 hr. = 60 min. = 3,600sec 1 day = 24 hrs. 1 wk. = 168 hrs 1 yr. = 365 da.	1 l. = 1.057 qts. 1 qt. = 0.9463 l. 1 l. = 2.113 pt. 1 oz = 29.57 ml 1 gal. = 3785 ml	1 cal. = 4.184 Joules

CONSTANTS	Prefix	Symbol	Exp. No.
R = 0.08206 L•atm/ mol.•K = 1.087 cal. / mol.•K = 8.314 J / mol.•K c = 3×10^8 m./sec. $R_H = 109,678 \text{ cm}^{-1}$ A = -2.18×10^{-18} J h = 6.63×10^{-34} J-sec $N_o = 6.02 \times 10^{23}$ units/mole	Tera- Giga- Mega- Kilo- Hecto- Deca- Deci- Centi- Milli- Micro- Nano- Pico-	T G M k h da d c m μ n p	10^{12} 10^9 10^6 10^3 10^2 10^1 10^{-1} 10^{-2} 10^{-3} 10^{-6} 10^{-9} 10^{-12}

ELECTRON CONFIGURATIONS:

The Aufbau Principle – Electrons normally occupy the lowest available energy level available.

The Aufbau Diagram – Order of filling of electron orbital structures.

Energy Level (n)	ORBITALS (l)				
	s	p	d	f	g
1	1s				
2	2s	2p			
3	3s	3p	3d		
4	4s	4p	4d	4f	
5	5s	5p	5d	4f	
6	6s	6p	6d		
7	7s	7p			