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Chemistry : Biology : Physics : Math

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COURSE FEATURES

This tutorial series is a carefully selected collection of core concept topics that cover the essential concepts. It consists of three parts:

1. Concept Tutorials – 24 essential topics
2. Problem-Solving Drills – 24 practice sets
3. Super Condense Cheat Sheets – 24 super review sheets

Core Tutorials

Self-contained tutorials...not an outline of information which would need to be supplemented by an instructor.

Concept map showing inter-connections of new concepts in this tutorial and those previously introduced.

Definition slides introduce terms as they are needed.

Visual representation of concepts.

Conceptual explanation of important properties and problem solving techniques

Animated examples of processes and cycles in human physiology.

A concise summary is given at the conclusion of the tutorial.

Problem Solving Drills

Each tutorial has an accompanying Problem Set with 10 problems covering the material presented in the tutorial. The problem set affords the opportunity to practice what has been learned.

Condensed Cheat Sheet

Each tutorial has a one-page cheat sheet that summarizes the key concepts and vocabularies and structures presented in the tutorial. Use the cheat sheet as a study guide after completing the tutorial to re-enforce concepts and again before an exam.

01: Introduction to DAT General Chemistry

Chapter Summary

This introductory tutorial is to warm you up on the DAT Chemistry course and get ready for the extensive coverage on required knowledge in the chemistry section of biological survey subtest. The DAT and CBT are defined and the scope of the chemistry section is introduced. The test-prep and test-taking strategies are also discussed. Finally, the tips in how to solve chemistry problems are described.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Dental Admissions Test
Computer based test format of the DAT
The DAT Natural Sciences focus areas
DAT General Chemistry comprehension areas
Test-preparation strategies
Test-taking strategies
How to attack and solve General Chemistry problems

Chapter Review

DAT - The DAT is the Dental College Admissions Test by American Dental Association (ADA). Its computer based test is a standardized exam that most US and Canadian dental schools require in order to consider a student for admission.

DAT General Chemistry - The science section of the DAT consists of 100 items that includes questions about: General Chemistry, Organic Chemistry and Biology. General Chemistry usually consists of about 30 items broken down as shown with the numbers representing the number of items usually tested within each focus area.

Test -Prep Strategies: (1) Smart Time Management (2) Key Concept Mastery (3) Practice-Practice-Practice (4) Systematic Review

Test-Taking Strategies: (1) Know the Test (2) Plan Your Attack (3) Build Focus (4) Zoom In On The Answer (5) Guess The Right Way

Chemistry Problem Solving: (1) Think of the answer before scanning the answer choices (2) Beware of the familiar (3) Aware of the Absolute (4) Deal with the Similar and Opposites (5) Know about the elimination and answer selection (6) Make educated guesses.

Chapter Summary

Chemistry, the study of matter and its interactions is introduced, along with descriptions of classic branches of chemistry (organic, inorganic, physical and analytical).

Matter is defined and categorized into pure substances (compounds or elements) and mixtures (homogeneous or heterogeneous). Energy is introduced, along with definitions of Potential Energy and Kinetic Energy.

Physical and chemical changes, and physical and chemical properties, are explained with hints as to how to classify a change or property.

The Metric and SI (International System) of units is used throughout chemistry. The metric system is based on prefixes showing the power of 10 used with base units describing the quantity measured.

Once careful measurements are taken, the precision with which they were measured can not be heightened while doing calculations, nor should it be lost and allowed to become less precise. Therefore, there are rules about counting "significant figures" which indicate which were measured for certain. Rules on performing calculations with significant figures will be introduced in a later tutorial.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Definition of Chemistry

Matter

- Pure Substances versus Mixtures
- Compounds versus Elements
- Homogeneous versus Heterogeneous

Energy

- Potential versus Kinetic

Changes

- Physical versus Chemical
- Possible signs of a Chemical Changes
- Common misconceptions between physical and chemical changes

Properties

- Physical versus Chemical
- Macroscopic versus Microscopic

The Metric and SI Systems

- Metric prefixes
- SI units

- Common non-SI units
- Significant Figures
- Why they're used
 - How to count them
- Fundamental Constants used in Chemistry

Chapter Review

Chemistry is the study of matter and its changes and interactions. Matter is anything that has mass and takes up space. There are two broad categories of matter—Pure Substances and Mixtures.

Pure Substances

Elements and compounds are both pure substances. A pure substance is when each particle is identical. Elements have each atom the same and compounds have each molecule the same. Compounds are atoms of more than one element chemically bonded together.

Mixtures

Mixtures are more than one type of pure substance physically mixed together. Mixtures can be categorized into homogeneous and heterogeneous. Homogeneous mixtures (also called solutions) look the same throughout; while heterogeneous have visible different types of matter.

Energy

Energy is the ability to produce heat or do work. There are two types of energy: Potential energy (or stored energy) and kinetic energy (energy due to motion).

Changes

Chemical changes produce a new substance while physical changes do not. Changes in state (melting, freezing, boiling, condensing, etc.) are physical changes. Dissolving is also a physical change, although it is often confused for a chemical change. Reacting with another type of matter, burning or rusting are examples of chemical changes. Often confused changes are melting (changing a solid to a liquid by adding heat), burning (chemically reacting with oxygen) and dissolving (combining two types of matter physically to produce a mixture). Mixtures can be separated by physical changes, compounds must be separated by chemical changes and elements cannot be separated by either.

Metric System

The metric system is based on prefixes that indicate a power of 10 with base units.

Metric Prefixes commonly used in chemist

Prefix	Symbc	Multiple
Kilo	k	1000
Deci	d	0.1
Centi	c	0.01
Milli	m	0.001
Micro	μ	0.000001
Nano	n	0.000000001

SI System

The International System of units gives a standard unit for each type of measurement.

SI Units commonly used in chemistry

Measurement	Unit	Symbol
Mass	Kilogram	kg

Volume	Liter	L
Temperature	Kelvin	K
Length	Meters	m
Time	Seconds	s
Amount of substance	Mole	Mol
Energy	Joule	J
Charge	Coulomb	C

There are also some important non-SI units as well.

Non-SI Units commonly used in chemistry

Measurement	Unit	Symbol
Length	Anstrom	Å
Pressure	Atmosphere	Atm
	Kilopascal	kPa
Energy	Calorie	cal
Temperature	Celcius	°C

Significant Figures

The significant figure rules are to allow people to read data or calculations and know with what precision the data was taken. The significant rules can be summarized in two rules: (1) If a decimal point is not present, count digits starting with the first the first non-zero number and ending with the last non-zero number; (2) If a decimal point is present anywhere in the number, start counting with the first non-zero number and continue until the end of the number. Rules on how to perform calculations with significant figures will be given in a future tutorial.

Fundamental Constants

Several numbers are used throughout chemistry and are important to be familiar with.

Fundamental constants commonly used in chemistry

Name	Symbol	Constant
Avogadro's #	N_A	$6.02 \times 10^{23} \text{ mol}^{-1}$
Speed of light	c	$3.0 \times 10^8 \text{ m/s}$
Gas constant	R	$8.31 \frac{\text{L} \cdot \text{kPa}}{\text{mole} \cdot \text{K}}$
		$0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mole} \cdot \text{K}}$
Planck's constant	h	$6.63 \times 10^{-34} \text{ J} \cdot \text{s}$
Charge of electron	e	$1.6 \times 10^{-19} \text{ C}$
Atomic mass unit	μ	$1.66 \times 10^{-24} \text{ g}$
Std Temp & Pressure	STP	273.15 K & 1 atm

Chapter Summary

Chemistry is the study of matter and its interactions. Matter is made of atoms, which make elements and molecules. Basic atomic structure and symbology is introduced in this tutorial, along with ions and isotopes.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Atoms

Ions

Element Symbols

Isotopes

Atoms, Elements, and Molecules

Chapter Review

Matter is made of atoms, which form elements and molecules.

Atoms

Atoms are the building block of matter. They are the smallest particle that retains the chemical properties of the element. Atoms are composed of protons, neutrons and electrons. Protons and neutrons are in the nucleus and contribute to the mass of the atom. Electrons are outside the nucleus, and together with the protons, contribute to the charge of the atom.

Ions

Ions are atoms that have a charge. There are an unequal number of positive protons and negative electrons.

Element Symbols

Element symbols are found on the periodic table. Element symbols can give lots of information: ${}^A_Z X^C$ where X is the symbol, A is the mass number (protons + neutrons), Z is the atomic number (protons), C is the charge (protons – electrons) and N is the number of atoms present.

Isotopes

Isotopes are elements of the same atom (same number of protons) with a different number of neutrons (and therefore a different mass). The atomic mass found on the periodic table is a

weighted average of all the isotope's individual masses. The mass number shown in the element symbol above refers only to 1 specific isotope.

Atoms, Elements and Molecules

Atoms are the smallest particle retaining the chemical properties of the element. Elements are pure substances that contain atoms with the same number of protons. Molecules are pure substances that contain more than one type of atom chemically bonded together.

Chapter Summary

It is critical to be able to write and read chemical formulas to function in a chemistry course. This tutorial introduces 4 different types of chemical formulas and gives instruction on how to write each type of formula.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Writing chemical formulas

- Binary Ionic
- Polyatomic Ionic
- Ionic with multivalent metals
- Binary Covalent
- Acids

Chapter Review:

Chemical formulas are written with rules according to the type of molecule.

Type 1: Binary Ionic

Between a metal and a non-metal and they'll end with "-ide" on the second word. The metal's symbol and charge are written, followed by the non-metals's symbol and charge. The charges are balanced by using coefficients to indicate the number of atoms.

Type 2: Polyatomic Ionic

They either begin with "ammonium" or end with "-ate" or "-ite" in the second word (except "hydroxide" and "cyanide"—those are polyatomic ions). A polyatomic ion is a group of atoms that together have a charge. The metal's symbol and charge are written first (or ammonium, NH_4^{+1} , the only polyatomic cation). The polyatomic anion's symbols and charge are written next. The charges are again balanced with subscripts. If a subscript is added to a polyatomic ion, use parenthesis around the ion.

Type 1 or 2 with Multivalent Metals

They will have roman numerals in the name. Multivalent metals are metals that have more than one possibility for the charge. The charge of the metal is indicated with roman numerals following the metal's name. The formula is then written following the rules for either Type 1 or Type 2.

Type 3: Binary Covalent

Between two non-metals. They will have use prefixes indicated the number of atoms.

"Mono-" is not used on the first element. The element symbols are written, and the prefixes indicate the subscript for each.

Type 4: Acids

The cation for an acid is H^+ . The anion is based on the format of the name: "hydro___ic acids" end with a single element; "___ic acids" end with the "___ate" polyatomic ion; "___ous acids" end with the "___ite" polyatomic ion. Write the correct anion's symbol and charge and then balance the charges with subscripts.

Chapter Summary

It is critical to be able to write and read chemical formulas to function in a chemistry course. This tutorial introduces 4 different types of chemical formulas and gives instruction on how to recognize and name each type.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Naming chemical formulas

- Binary Ionic
- Polyatomic Ionic
- Ionic with multivalent metals
- Binary Covalent
- Acids

Chapter Review

Chemical formulas are named with rules according to the type of molecule.

Type 1: Binary Ionic

Between a metal and a non-metal—one element of each type. Write the name of the first element (the metal) and then the name of the second element with “-ide” replacing the last syllable. Subscripts are not important when using this type

Type 2: Polyatomic Ionic

It will contain more than 2 elements—with at least one being a metal and one being a non-metal. Write the name of the first element or polyatomic ion. Write the name of the second element or polyatomic ion. If the anion is an element, change the ending to “-ide”; if the anion is a polyatomic ion, do not change the ending. Polyatomic ions must match exactly—including the subscripts. If there is parenthesis, the polyatomic ion is inside the parenthesis.

Type 1 or 2 with Multivalent Metals

They'll start with Co, Cr, Cu, Fe, Hg, Pb, Sn. Multivalent metals are metals that have more than one possibility for the charge. The name of the metal is written, followed by roman numerals in parenthesis indicating the charge of the metal. The charge is determined by knowing the charge of the anion and knowing that the overall charge of the molecule is 0. The name of the anion is written—changing the ending of a single element anion to “-ide.”

Type 3: Binary Covalent

Between two non-metals. Write the name of the first element with a prefix indicating the subscript (do not use "mono-" with the first element). Write the name of the second element with a prefix indicating the subscript and "-ide" as the ending syllable.

Type 4: Acids

The cation for an acid is H^+ . The name is based on the anion. A single element anion is named as "hydro____ic acid"; a "____ate" polyatomic anion is named as "____ic acid"; a "____ite" polyatomic anion is named as "____ous" acid.

Chapter Summary

Chemical reactions are the “sentences” of chemistry that show what molecules enter a chemical change and what molecules are produced in the change. The tutorial will introduce the components of a chemical reaction, the common types of chemical reactions and how to predict products of simple chemical reactions

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Components of a chemical reaction

Common types of chemical reactions

- Composition
- Decomposition
- Single replacement
- Double replacement
- Neutralization
- Precipitation
- Redox

Oxidation numbers

- Determining
- Using to determine redox reactions

Determine products of a double replacement reaction

Using solubility rules to determine a precipitate

Writing net ionic reaction

Chapter Review:

Chemical reactions are the “sentence” that shows a chemical change.

Components of a chemical reaction

Chemical reactions are made of reactants are listed first, followed by an arrow that indicates “yields,” “produces,” or “forms.” The arrow is followed by the products of the chemical reaction. Chemical reactions can also show states of matter and energy changes.

Common types of chemical reactions

There are several common types of chemical reactions. Composition reactions are when more than one type of matter combines to form one molecule. Decomposition reactions are the opposite of composition reactions. Single replacement reactions involve an element reacting with a compound and replacing one of the elements in the compound. A double replacement

reaction has two compounds that switch ions. A neutralization reaction is a double replacement reaction between an acid and a base. Precipitation reactions are double replacement reactions that produce an insoluble compound. Redox reactions involve the transfer of electrons from one atom to another, resulting in the change of an oxidation number.

Oxidation numbers & Redox reactions

Oxidation numbers are the charge on an atom if the electrons involved in the bond are assigned to the more electronegative atom in the bond. When oxidation numbers change during a chemical reaction, it is a redox reaction.

Double replacement reactions and precipitations

Solubility rules can be used to determine if a double replacement reaction forms a precipitate, an insoluble ionic compound.

Net Ionic Reactions

Net ionic reactions remove all spectator ions (ions that dissociate on both sides of the reaction and remain unchanged).

Chapter Summary

The Law of Conservation of Matter says that matter cannot be created nor destroyed. That means that the atoms that are on the reactant side also must appear on the products side. Therefore, reactions need to be balanced. This tutorial shows methods that can be used to balance simple reactions as well as more complicated redox reactions.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Using The Law of Conservation of Mass to balance the equations
Choosing which atom to begin balancing with
Inspection method of balancing
Oxidation number method of balancing
Half-reaction method of balancing

Chapter Review

Chemical reactions are balanced with coefficients until the numbers of each atom are equal on the left and the right.

The Law of Conservation of Mass

The Law of Conservation of Mass states that the mass of the reactants equal the mass of the products. Atoms cannot be created nor destroyed in a chemical change—therefore, the number of each type of atom on each side of the reaction must be equal. Coefficients are used to balance chemical reactions.

Choosing which atom to start with

Start with elements that appear only one time on each side and elements that are in the most complex compounds. End with elements that appear more than once on a side or elements that appear uncombined on one side or the other.

Inspection method of balancing

The Inspection Method is used to balance the simplest reactions. It includes:

1. Make a list of the elements in the reaction
2. Count the number of each type of atom on each side of the reaction
3. Add coefficients to balance the number of atoms
4. Determine the total charge of each side of the reaction and use coefficients to balance charge.

5. When all elements and charge are balanced, place a "1" in any empty coefficient location to indicate that you're done.

Oxidation number method of balancing

The oxidation number method is used for balancing simple redox reactions that cannot be easily balanced by the inspection method. It includes:

1. Determine the oxidation numbers of each atom.
2. Determine the net change in charge. Use the net change to determine the ratio of atoms that would cancel out the net charge change.
3. Use the ratio as coefficients in the simplest compounds containing those elements.
4. Finish balancing by the inspection method.

Half-reaction method of balancing

The half-reaction method is for the most difficult redox reactions:

1. Use oxidation numbers to determine what's oxidized and what's reduced.
2. Write two half-reactions, one for reduction and one for oxidation
3. Balance all elements except H and O using inspection method.
4. **For an acid redox reaction:** Balance the O's by adding H₂O to the side needing more O. **For a base redox reaction:** Balance O by adding twice as many OH⁻ to the side needing more O.
5. **For an acid redox reaction:** Balance the H's by adding H⁺ to the side needing more H's. **For a base redox reaction:** Balance H's by adding H₂O to the side needing more H's.
6. Determine the charge of each side of each reaction. Balance the charges by adding electrons to the side with the more positive charge for each reaction.
7. Multiply the half-reactions by factors that will allow the electrons to cancel out.
8. Add the two half-reactions back together.
9. Cross out anything that appears the same on both sides.

Chapter Summary

Atoms and molecules are too small to be counted individually, so a counting unit of “mole” is used. Molar masses are used to convert between mass and moles of atoms or molecules.

There are several ways to express the ratio of moles of atoms in a sample. The tutorial introduces percent composition, empirical formula and molecular formula.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

The Mole

- Mole relationships

Molar Mass

- Calculating
- Using in mole conversions

Percent Composition

Empirical formulas

Molecular formulas

Chapter Review

Atoms and molecules are so tiny that they need to be counted in very large quantities in order to result in a “workable” number.

The mole

The mole (abbreviation: mol) is used to count particles. 1 mole of anything is 6.02×10^{23} of that thing.

Molar mass

We cannot count out a mole of particles in the lab, therefore a connection between moles and mass is needed. The molar mass is the mass (in grams) for 1 mole of the particle. Molar mass (also often called Molecular mass, formula weight, etc.) is found by adding the atomic masses (multiplied by the number of that atom) for each atom in the molecule. Molar mass can then be used in dimensional analysis conversions as the equality between grams and moles. The molar mass value always goes with the “grams” in such conversions.

Percent Composition

“Percent” is $(\text{part} / \text{whole}) \times 100$. The “part” is the mass of an individual element while the “whole” is the mass of the whole molecule or sample. These calculations can be done with lab

data—grams of both the individual element and the whole sample—or with chemical formulas. In the case of chemical formulas, the molar mass is used in the calculations.

Empirical Formula

The empirical formula is the lowest possible ratio of atoms in a molecule. The ratio of atoms is the same as the ratio of moles of atoms. The mass of each element is converted to moles and then all the moles are divided by the smallest to reach the lowest possible whole number ratio. That ratio is used as the subscripts in writing the empirical formula.

Molecular Formula

The molecular formula is the actual ratio of atoms in a molecule. The molar mass of the empirical formula is compared to the given molar mass of the molecular formula. The comparison results in a factor that is used to multiply the subscripts of the empirical formula to reach the subscripts of the molecular formula.

Chapter Summary

Stoichiometry uses the principles of dimensional analysis to use information about one species in a chemical reaction to determine information about a different species in the same reaction. This tutorial guides you through stoichiometric calculations involving moles, mass, solution volume and gas volume.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Review

- Chemical equations
- Dimensional Analysis
- KUDOS method
- Molarity

Stoichiometry

- Mole-Mole problems
- Mole-mass problems
- Mass-mass problems
- Mole-volume problems for solutions
- Mole-volume problems for gases

Chapter Review

Stoichiometry is the culmination of many concepts in chemistry. It integrates dimensional analysis, moles and molar mass, molar volumes of gases, concentrations and balanced chemical equations to use information about one species in a reaction to determine information about another species in the same reaction.

Review

Chemical equations are used to relate reactants and products in a chemical reaction. Dimensional analysis is a technique that uses ratios of equivalents to convert units. The KUDOS method is a method for solving word problems. Molarity is a concentration unit showing the ratio of moles of solute to volume of solution.

Mole-Mole problems

The balanced equation gives the ratio of moles of each species in the reaction. The balanced equation is used to determine mole ratios for use in dimensional analysis when converting from the moles of one species to the moles of another. Moles is the only unit that may be

used to convert from one chemical to another—mass or volume cannot be used to convert between chemicals, just within the same one.

Mole-Mass & Mass-Mass problems

Molar mass is an equivalent between the mass in grams and moles of a substance. The molar mass is used to convert between moles of a substance and mass in a stoichiometry problem.

Mole-Volume problems for solutions

The concentration, in molarity, of a solution gives the equivalent for moles of the solute and liters of solution. The molarity can be used to convert between moles and volume of a solution in stoichiometry problems.

Mole-Volume problems for gases

The volume of 1 mole of any gas at standard temperature and pressure is 22.4 L. The molar volume of a gas can be used to convert between moles and volume of a gas at STP in stoichiometry problems. If a problem asks for the volume at non-standard temperature or pressure, find the STP volume using stoichiometry and then use the appropriate gas law to convert to the desired temperature or pressure.

Chapter Summary

The most energetic of the three states of matter, gases, has several laws that describe behavior. The Kinetic Molecular Theory attempts to explain gas behavior. This tutorial will introduce the KTM, pressure, gas behavior, several gas laws, effusion and diffusion.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Pressure

- What is pressure?
- Atmospheric pressure
- Units of pressure

Kinetic Molecular Theory

- Ideal versus real gases

Explanations of gas behavior

- Relationships between pressure, volume and temperature
- Internal versus external pressure

Gas laws

- Avogadro's
- Boyles'
- Charles'
- Combined
- Partial Pressure
 - Mole fractions
- Ideal
 - With density
 - With molar mass
- Real

Gas stoichiometry

Effusion & Diffusion

Chapter Review

Gases are the state of matter with the greatest amount of energy.

Pressure

Pressure is created by gas particles running into the wall of the container. Pressure is measured in many units: 1 atm = 101300 Pa = 101.3 kPa = 760 mm Hg = 14.7 psi. Atmospheric pressure is the pressure due to the layers of atmosphere above us.

Kinetic Molecular Theory

The Kinetic Molecular Theory has several assumptions for ideal gases.

- Gases are made of atoms or molecules
- Gas particles are in rapid, random, constant motion
- The temperature is proportional to the average kinetic energy
- Gas particles are not attracted nor repelled from each other
- All gas particle collisions are perfectly elastic (they leave with the same energy they collided with)
- The volume of gas particles is so small compared to the space between them that the volume of the particle is insignificant

Real gases do have a volume (that takes up space which other particles cannot occupy) and they do have attractions/repulsions from one another as well as in-elastic collisions.

The KMT is used to understand gas behavior. Pressure and volume are inversely proportional. Pressure and temperature are directly proportional. Pressure and number of particles are directly proportional.

An expandable container will expand or contract so that the internal and external pressures are the same. Non-expandable containers will explode or implode if the difference in the pressures is too great for the container to withstand.

Gas Laws

Symbols for all gas Laws:

P = Pressure; V = Volume; n = moles; T = Temperature (in Kelvin);

R = Gas constant $8.31 \frac{L \times kPa}{mole \times K}$ or $0.0821 \frac{L \times atm}{mole \times K}$

"a" and "b" = correction factors for real gases

Combined Gas Law: $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$ When something is held constant, it cancels out.

Dalton's Law of Partial Pressure: $P_{total} = \sum P_{of\ each\ gas}$

Mole fraction: $\chi_A = \frac{mole_A}{mole_{total}}$

Partial Pressure and mole fraction: $P_A = \chi_A P_{total}$

Ideal Gas Law: $PV = nRT$

Ideal Gas Law with Molar Mass: $PV = \frac{m}{MM} RT$

Ideal Gas Law with Density: $P = D \frac{RT}{MM}$

Real Gas Law: $\left(P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$

Gas Stoichiometry

Use the molar volume of a gas at STP (1 mole of any gas at STP = 22.4 L) to convert between moles and liters of a gas in stoichiometry. Then use the appropriate gas law to find the volume at non-STP conditions.

Diffusion and Effusion

Diffusion is the rate at which a gas travels through a container. Effusion is the rate at which gas escapes through a tiny hole in the container. Both are inversely proportional to the square root of the molar mass (heavier molecules travel slower). **Graham's Law:**

$$\frac{r_1}{r_2} = \sqrt{\frac{MM_2}{MM_1}}$$

Chapter Summary

Liquids and solids are condensed states of matter. This tutorial introduces the intermolecular forces between molecules in the condensed states. Properties, such as Vapor Pressure, are explained. The various types of solid structure are shown. Changes in state, and the energy associated with them, are also discussed.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Intermolecular Forces

- London Dispersion Forces
- Dipole-Dipole and Ion-Dipole forces
- Hydrogen bonding

Properties of a liquid

Vapor pressure

Properties of solids

Solid structures

- Amorphous solids
- Atomic solids
 - Metallic
 - Network
- Molecular solids
- Ionic Solids

Phase change

- Terms
- Equilibriums
- Energy associated with phase changes

Chapter Review

Liquids and solids are condensed states of matter that have intermolecular forces.

Intermolecular forces

Physical attractions between separate molecules are intermolecular forces. The weakest of these forces is present in all molecules, **London Dispersion Forces**. It's due to the temporary ganging up of electrons on one side of the molecule. This creates a temporary dipole that can be attracted to other dipoles, temporary or permanent. Molecules with a permanent dipole can also display **Dipole-Dipole** attractions, or **Ion-Dipole** attractions with

an ion. Since the ability to form these attractions isn't temporary, as with London Dispersion Forces, dipole attractions are stronger. When a hydrogen atom is bonded to a very electronegative atom, N, O or F, it forms a very strong dipole. This extra strong dipole can form strong attractions with an N, O or F on another molecule, called **Hydrogen Bonding**, which is the strongest IMF.

Properties of liquids and solids

Liquids have definite volumes, but not definite shapes, the particles are free to move past each other and they are not very compressible. Solids have definite volumes and shapes, the particles are not able to move past each other and they are not compressible at all.

Vapor Pressure

Vapor pressure is created when molecules on the surface of a solid or liquid have enough energy to escape the intermolecular forces and become a gas. Once it is a gas, it can create pressure. As temperature increases, more molecules have the minimum energy to evaporate, and vapor pressure increases. At first, many molecules escape and the volume of the solid or liquid decreases. But after a while, some begin to collide with the surface of the liquid and rejoin the liquid again. Eventually equilibrium is established.

Solids Structures

Amorphous solids have no repeating structure. Crystalline solids do have a pattern of repeating units. Atomic solids have atoms as the repeating unit and can be metallic (electrons are shared in a large pool throughout the metal) or network (where each atom is covalently bonded to each other atom) solids. Molecular solids have molecules that are all covalently bonded to one another. Ionic solids have a network of ions arranged to maximize electrostatic attractions and minimize repulsions.

Phase Changes

During melting or boiling, intermolecular forces are broken. During freezing or condensing, intermolecular forces are formed. The melting/freezing point is when the vapor pressure of the solid and liquid are equal and the two states of matter are at equilibrium. The boiling/condensing point is when the vapor pressure of the liquid is equal to the atmospheric pressure and is the temperature when the two states of matter are at equilibrium.

Chapter Summary

Solutions are common in chemistry. This tutorial introduces how solutions are formed, the factors that affect solubility, and several ways in which concentration is expressed, electrolyte solutions, colligative properties and colloids. It also reviews using concentration in stoichiometric calculations.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Process of solution formation

Factors affecting solubility

Concentration calculations

- Expressing concentration
- Dilution calculations
- Using concentration in stoichiometry

Solubility equilibrium

Chapter Review:

Solutions are formed when a solute is dissolved in a solvent.

Forming solutions

In order for a solution to form, the solute intermolecular forces must be broken as well as the solvent intermolecular forces. Then the solute and solvent form new intermolecular forces with each other. If the energy required to break the intermolecular forces is much greater than the energy released when the new forces are formed, the solution will not form.

Factors affecting solubility

For gases, as the pressure of the gas above the solution increases, the solubility of the gas increases. For gases, as the temperature of the solution increases, the solubility of the gas decreases. For most solids, as temperature increases, the solubility increases.

Concentration calculations

There are many ways to express concentration (which is the ratio of solute to solvent or solution).

% by mass: $\% \text{ mass} = \frac{\text{mass solute}}{\text{mass solution}} \times 100$

The mass units must match!

% by volume: $\% \text{ volume} = \frac{\text{volume solute}}{\text{volume solution}} \times 100$

The volume units must match!

% mass/volume: $\% \text{ mass / volume} = \frac{\text{mass solute}}{\text{volume solution}} \times 100$

The volume unit is mL

Molarity (M): $\text{Molarity} = \frac{\text{moles solute}}{\text{L solution}}$

Molality (m): $\text{Molality} = \frac{\text{moles solute}}{\text{kg solvent}}$

A sample becomes diluted (less concentrated) when more solvent is added. The dilution equation is

$M_1V_1 = M_2V_2$ M_1 = original molarity V_1 = original volume M_2 = new molarity
 V_2 = new volume. Volume units must match!

Solubility equilibrium

The solubility equilibrium constant is written for a dissolution reaction (a solid compound dissociates when dissolved in water). The pure solid reactant is not included in the expression. A saturated solution is one that is at equilibrium.

Chapter Summary

Many reactions in chemistry involve acids and bases. This tutorial will introduce the three definitions for acids and bases, as well as discussing the concept of strong versus weak acids and bases. Conjugate acids and bases will be discussed, as well as the acid or base properties of the salts produced. Equilibrium constants will be used for acid/base reactions as well as for the autoionization of water. The pH scale and pH calculations will be introduced, along with buffers and titrations.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Acids and base definitions

- Arrhenius
- Bronsted-Lowry
- Lewis

Properties of acids and bases

Strong versus weak acids and bases

Conjugates of acids and bases

Equilibrium of acids and bases

pH scale

Acid and base properties of salts

Buffers

Titration

Chapter Review

Acids and bases are used throughout chemistry.

Definitions of acids and bases

There are three common definitions of acids and bases:

- **Arrhenius acid:** Produces hydronium ion in water.
- **Arrhenius base:** Produces hydroxide ion in water.
- **Brønsted-Lowry acid:** Donates a proton (H^+)
- **Brønsted-Lowry base:** Accepts a proton
- **Lewis acid:** Accepts electrons
- **Lewis base:** Donates electrons

Strong acids and bases are ones in which most molecules perform their “duty” while weak acids and bases only have a few acid and base molecules that act as acids and bases. There are only a few strong acids and bases to remember—the rest will most likely be weak. Strong

acids: HCl, HBr, HI, HNO₃, HClO₃, HClO₄. Strong bases: NaOH, KOH, Ca(OH)₂, Ba(OH)₂, Sr(OH)₂

A conjugate acid is what remains after a base does its "job". A conjugate base is what's left after an acid does its "job." Strong acids or bases form a weak conjugate and vice versa.

Equilibrium of acids and bases

Equilibrium constants can be written for acid and base dissociation reactions. Water also autoionizes to form hydrogen and hydroxide ions. The equilibrium constant for the acid dissociation reaction \times the base dissociation constant for the conjugate base = the water dissociation constant at that temperature.

pH

The pH scale is a logarithmic scale to measure the acidity of a solution. $pH = -\log[H_3O^{+}]$

Strong acids and bases can be assumed to dissociate completely. Therefore, the concentration of the original strong acid or base is assumed to equal the concentration of the hydrogen or hydroxide ion. For weak acids or bases, the equilibrium constant and ICE charts are used to determine the concentration of the hydronium ion before solving for pH.

Acid/base properties of salts

Some salts can have acid/base properties based on the acid or base they are based off of.

Salts from

- Weak acid + strong base = Basic
- Weak acid + weak base = Neutral
- Strong acid + weak base = Acidic
- Strong acid + Strong base = Neutral

Buffers

A buffer is a solution containing a weak acid and its conjugate base or a weak base and its conjugate acid. Buffers resist changes in pH when acids or bases are added. The pH will still change, but much less than if it was plain water. Buffers use the Henderson-Hasselbach

equation: $pH = pKa + \log\left(\frac{[base]}{[acid]}\right)$

Titrations

Titrations are used to find the concentration of an unknown solution using a solution of known concentration. An indicator is used that changes color at the stoichiometric point (the point at which no reactants are left over) based on the pH of the products that are in solution at that point. Stoichiometry is used at that point to determine the unknown concentration.

Chapter Summary

Equilibrium is established when the rate of the forward reaction is equal to the rate of reaction of the reverse reaction in a reversible reaction. This tutorial introduces how equilibrium is established, how equilibrium constant expressions are written and how equilibrium constants are calculated. A problem solving technique for equilibrium problems will be illustrated in the examples. The use of the reaction quotient to determine if a system is at equilibrium is illustrated. Le Chatelier's principle will be introduced, along with solubility equilibrium.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Dynamic equilibrium

Equilibrium constant

- Writing expressions
- Calculating constants
- Meaning of the equilibrium constant

Reaction quotient

- Writing expressions and finding the value
- Using it to determine if it's at equilibrium

Solving equilibrium problems

- Using the ICE chart

Solubility equilibrium

Le Chatelier's principle

Chapter Review

Dynamic equilibrium is when the rate of the forward and reverse reactions are equal. The reaction will appear to have stopped, as the concentrations of each species won't change, but the reaction continues to proceed in both directions.

Equilibrium constants

Equilibrium constant expressions are the ratio of the concentrations of products to reactants using their coefficients from the balanced equation as exponents. Pure solids and liquids are not included in the equilibrium constant expression. Equilibrium concentrations are plugged into the expression to solve for the equilibrium constant. A large equilibrium constant means that the reaction "lies to the right" with more products than reactants at equilibrium. A small constant means that the reaction "lies to the left" with more reactants the products present at equilibrium.

Reaction Quotients

The reaction quotient is the value when concentrations at any time are plugged into the equilibrium constant expression. If the reaction quotient equals the equilibrium constant, then the system is at equilibrium. If the reaction quotient is too large, the reaction will proceed to the left to produce more reactants (and reduce products). The opposite will happen if the reaction quotient is too small.

Solving equilibrium problems

A simple technique for solving equilibrium problems is the ICE chart method:

- Make a table with the reactants and products across the top
- Place "ICE" down the left hand side, for Initial, Change and Equilibrium.
- Fill in any given information from the problem.
- Use the balanced equation's stoichiometric ratio to determine the "change" row.

To solve the problem using the ICE chart:

- If an equilibrium concentration is known, you can determine the "change" and find the other equilibrium concentrations to plug in and solve for K
- If you don't know any equilibrium concentrations, write expressions for them and plug in the expressions into your "K" equation to solve for them.
- If the K is very tiny (10^{-5} or smaller), you may approximate that the change is insignificant compared to the original value (if the original value is > 0) e.g.: $0.25 \text{ M} - x \approx 0.25 \text{ M}$

Solubility equilibrium

The solubility equilibrium constant is written for a dissolution reaction (a solid compound dissociates when dissolved in water). The pure solid reactant is not included in the expression. A saturated solution is one that is at equilibrium.

Le Chatelier's Principle

Le Chatelier's principle says that if a system at equilibrium is stressed or changed, the system will shift to reach equilibrium again. Whatever you do to the system, it will un-do it. For example, if you add more reactants, the reaction will proceed to the right to get rid of the extra reactants and make more products.

Chapter Summary

Thermodynamics is the study of heat changes during processes. The heat changes of physical processes will be introduced using specific heat capacity, heats of fusion and vaporization, and calorimetry. Heat changes of a chemical reaction will be discussed with heat of formation, Hess's Law and calorimetry.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Energy, temperature and heat

Energy and physical changes

- Specific heat capacity
- Calorimetry
- Heat of fusion
- Heat of vaporization
- Heating curves

Energy and chemical changes

- Heat of formation
- Hess's Law
- Heat of reaction

Entropy

Free Energy

- And equilibrium

Chapter Review

Physical and chemical changes occur with changes in heat.

Energy, heat and temperature

Several definitions are useful in understanding thermodynamics:

- **Energy:** The ability to do work or supply heat.
- **System:** Particles under-going change.
- **Surroundings:** Everything surrounding the system.
- **Temperature:** proportional to the average kinetic energy of the particles.
- **Heat (q):** Flow of energy from a hotter object to a cooler object.
- **Enthalpy (H):** Takes into account internal energy, pressure and volume. In an open-air environment, it's the same as heat.

Physical changes

The specific heat capacity is the amount of energy that can be absorbed before temperature begins to change. $\Delta H = m \times C_p \times \Delta T$ where m = mass; C_p = specific heat capacity and $\Delta T = T_2 - T_1$. During changes in state, the heat of fusion and heat of vaporization are used.

Melting: $\Delta H = m \times H_{fus}$ H_{fus} = enthalpy of fusion. Boiling: $\Delta H = m \times H_{vap}$ H_{vap} = enthalpy of vaporization. Heating curves show the combination of the processes—changing temperatures and changing states.

Chemical changes

The heat of formation is the energy change when a compound is formed from its elements. Hess's law says that since energy is a state function, the path doesn't matter—only where you began and ended, then the heat of a reaction can be found by adding up stepwise reactions that add up to the overall chemical reaction. This allows the heat of reaction to be found from formation values: $\Delta H_{rxn} = \sum H_f \text{ prod} - \sum H_f \text{ react}$

Calorimetry

When the system loses energy, the surroundings gain it and vice versa. Therefore, the energy change in the surrounding can be measured and used to determine information about the system. Calorimetry can be used for physical or chemical processes.

Enthalpy, Entropy and Free energy

Entropy is disorder or randomness. All spontaneous processes result in a net increase in entropy for the universe. The spontaneity of a process (shown with a negative free energy value) is found by relating enthalpy, entropy and temperature to find free energy.

$$\Delta G = \Delta H - T\Delta S$$

Enthalpy	Entropy	Spontaneous at
+	+	High temps
+	-	Never
-	-	Low temps
-	+	All temps

Chapter Summary:

Kinetics is the study of reaction rates. This tutorial introduces factors affecting the rate of reaction, reaction mechanisms, writing rate laws (both differential and integrated), half-lives, and relating the rate law constant to the minimum energy needed for reaction (the activation energy).

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Collision Theory

Reaction Coordinate Diagrams

Factors affecting rate

Reaction mechanisms

- Elementary steps

Rate laws

- Differential
- Integrated

Half-life

Relationship of rate law constant to activation energy

Chapter Review

Kinetics is the study of the rate of reaction and the factors that affect it.

Collision theory

In order for a reaction to occur, the molecules must collide in the correct orientation with the minimum energy needed for the transition from the reactants to the products (the activation energy). Only a very few collisions meet these requirements and result in a reaction.

Reaction Coordinate Diagrams

Reaction coordinate diagrams show the energy of the reactants, the activation energy up to the activated complex, or transition state (the in-between state between the reactants and the products), and the energy of the products. The overall energy change of the reaction is also shown.

Factors affecting rate

Increasing the temperature increases the number of collisions, and also the number of collisions with the needed energy. Therefore, increasing temperature increases the rate of reaction. Increasing the concentration or the surface area also increases the number of collisions, therefore increasing the chance that a successful collision will occur—which

increases rate. Adding a catalyst, a species that increases the rate of reaction without being used up in the reaction, also increases the rate.

Reaction mechanisms

Reaction mechanisms are a set of elementary steps. Each elementary step shows which molecules must collide at one time in order to produce a reaction. The elementary steps add up to the overall chemical reaction. The slowest elementary step is the rate-determining step. The reaction rate law can be written from the correct rate-determining elementary step—but it cannot be written from the overall chemical reaction. One way of evaluating the possibility of a proposed reaction mechanism is to see if it matches the experimentally found rate law.

Rate laws

Differential rate laws relate the rate of reaction to the concentration of the reactants. Each reactant's concentration is taken to a power, or "order", that corresponds to the number of that species that must collide in the rate-determining step. The rate law has a rate constant that is different for each reaction at each temperature. Integrated rate laws relate the concentration of a species over time. If one rate law is known, the other rate law can be found—they come in "matched" sets. The half-life (time that it takes for half of the reactants to react away) can be found using the integrated rate law and setting the $[A]$ at $t_{1/2}$ to $\frac{1}{2}[A]_0$.

Rate law constants and activation energy

The higher the activation energy, the less often a collision will result in a successful reaction. Therefore, the higher the activation energy, the lower the temperature. The Arrhenius equation relates the rate law constant to the activation energy at a given temperature.

Chapter Summary

Electrochemistry is the study of the inter-change between chemical and electrical energy. This tutorial begins with a review of oxidation states, redox reactions, and balancing redox reactions from earlier tutorials. Voltaic cells, line notation standard cell potential and electromotive force (cell potential) are explained, along with electrolytic cells. Electrochemistry is then related to free energy and equilibrium.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Review

- Oxidation states
- Redox reaction
- Balancing redox reactions

Voltaic cells

- Composition

Cell potentials

- Standard reduction potential
- Voltaic cell potential

Electrolytic cells

Electrochemistry and free energy

Electrochemistry and equilibrium

Chapter Review

Electrochemistry studies how chemical and electrical energy are converted.

Oxidation numbers & Redox reactions

Oxidation numbers are the charge on an atom if the electrons involved in the bond are assigned to the more electronegative atom in the bond. When oxidation numbers change during a chemical reaction, it is a redox reaction.

Oxidation number method of balancing

The oxidation number method is used for balancing simple redox reactions that cannot be easily balanced by the inspection method. It includes:

1. Determine the oxidation numbers of each atom.
2. Determine the net change in charge. Use the net change to determine the ratio of atoms that would cancel out the net charge change.
3. Use the ratio as coefficients in the simplest compounds containing those elements.
4. Finish balancing by the inspection method.

Half-reaction method of balancing

The half-reaction method is for the most difficult redox reactions:

1. Use oxidation numbers to determine what's oxidized and what's reduced.
2. Write two half-reactions, one for reduction and one for oxidation
3. Balance all elements except H and O using inspection method.
4. **For an acid redox reaction:** Balance the O's by adding H₂O to the side needing more O. **For a base redox reaction:** Balance O by adding twice as many OH⁻ to the side needing more O.
5. **For an acid redox reaction:** Balance the H's by adding H⁺ to the side needing more H's. For a base redox reaction: Balance H's by adding H₂O to the side needing more H's.
6. Determine the charge of each side of each reaction. Balance the charges by adding electrons to the side with the more positive charge for each reaction.
7. Multiply the half-reactions by factors that will allow the electrons to cancel out.
8. Add the two half-reactions back together.

Voltaic cells

A voltaic cell separates the reduction and oxidation reaction and forces the electrons to flow over a wire (producing electricity) from the oxidation reaction (at the anode) to the reduction reaction (at the cathode). The cell consists of the two separate half reactions, metal electrodes and a wire for conducting the electrons, and a salt bridge for balancing the charge build-up to extend the time the cell will operate. Line notation is a short-hand way of describing a cell:

- Anode written first
- Reactants written 1st on each side
- Anode & Cathode separated with ||
- Different states of matter within same side separated with |
- Same states of matter within same side separated with a comma

Cell potentials

The cell potential (or electromotive force) of a voltaic cell is due to the potential energy difference of the electrons before the transfer and after the transfer. A standard reduction potential is the potential that would be produced between a given half-reaction and hydrogen (hydrogen's standard reduction potential has been defined as 0). The standard reduction potentials can be used to calculate the cell potential: $EMF = \text{cathode} - \text{anode}$. Positive EMF values indicate a spontaneous process.

Electrolytic cells

An electrolytic cell is the opposite of a voltaic cell. An electrolytic cell converts electrical energy into chemical energy by forcing a reaction to proceed in the non-spontaneous direction by putting electricity in. The voltage need to force the reaction in the opposite direction is at least that produced by the spontaneous process.

Electrochemistry and free energy

The free energy of a system can be defined as the amount of work that can be done by the system. The flow of electrons can do work. Therefore, the free energy of the system can be defined as: $\Delta G = -n \times F \times EMF$ $\Delta G = \text{free energy (in J)}$; $n = \# \text{ of moles of electrons transferred}$; $F = 1 \text{ Faraday}$; $EMF = \text{cell potential}$.

Electrochemistry and equilibrium

When a cell reaches equilibrium, the cell stops reacting.
The Nernst equation relates EMF to equilibrium:

$EMF = EMF^\circ - \frac{RT}{nF} \ln Q$ EMF = cell potential at current conditions; EMF° = cell potential at standard state (1 atm & 25°C); $R = 8.31 \text{ J/mole}\times\text{K}$; T = temperature (in Kelvin); n = moles electrons transferred; $F = 1 \text{ Faraday}$; Q = reaction quotient

A cell stops when it reaches equilibrium. At equilibrium, $EMF = 0$ and $Q = K$

$\ln K = \frac{n \times F \times EMF^\circ}{RT}$ where K = equilibrium constant

Chapter Summary

Atoms are composed of 3 subatomic particles: protons, neutrons and electrons. This tutorial review the basic atomic structure information presented in an earlier tutorial and introduces the concept of placing electrons in energy levels, subshells and orbitals to show an electron configuration.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Atomic structure

- 3 subatomic particles
- Determining number of protons and electrons in an atom or ion

Electron configurations

- Energy levels, subshells and orbitals
- Rules for writing electron configurations
- Boxes & Arrows configuration
- Spectroscopic configuration
- Noble gas configuration
- Exceptions to the rules

Quantum numbers

- Assigning numbers to an electron
- Identifying a non-possible quantum number

Chapter Review

Electrons are involved in bonding, and therefore, their arrangement in an atom is very important.

Subatomic particle

Atoms are composed of protons, neutrons and electrons. The atomic number is equal to the number of protons. The charge of the atom or ion is equal to the number of protons – the number of electrons. The number of electrons can be determined knowing the atomic number and charge.

Energy levels, subshells and orbitals

Electrons are placed in energy levels. These energy levels are sub-divided into subshells (labeled s, p, d or f). The s subshell is the lowest energy and begins in level 1. The p subshell is higher energy and therefore doesn't begin until level 2. The d is higher energy and begins in level 3 and the f is even higher energy and begins in level 4. The subshells are

further sub-divided into orbitals (s has 1 orbital, p has 3 orbitals, d has 5 orbitals and f has 7 orbitals). Each orbital can hold 2 electrons.

Rules for writing electron configurations

The **Aufbau principle** states that energy levels must be filled from the lowest to the highest and you may not move on to the next level unless the previous level is full. Use the periodic table as a guide (read left to right):

1			
2			2p
3			3p
4		3d	4p
5		4d	5p
6	4f	5d	6p
7	5f	6d	7p

Hund's Rule says that when placing electrons in orbitals of equal energy, place one in each orbital before doubling up in order to arrive at the lowest energy configuration. The **Pauli Exclusion Principle** states that when electrons do share an orbital, they must be of different "spin."

Writing electron configurations

The **Boxes and Arrows method** uses boxes to show orbitals and arrows to signify electrons. An up arrow and a down arrow have different "spins." The **spectroscopic method** uses superscripts to show the number of electrons in a subshell (specific orbitals are not shown). The **noble gas method** uses a noble gas (the far right column) to represent the inner, or core, electrons and just shows the outer level of electrons using the same method as spectroscopic.

Exceptions to the rules

There are a few exceptions to the rules listed above when filling electron configurations. A half-full "s" orbital and a "d" subshell with 5 or 10 is more stable than following the Aufbau Principle. **Cr, Mo, W:** $s^1 d^5$ and **Cu, Ag, Au:** $s^1 d^{10}$

Quantum numbers

Quantum numbers are used to describe the location of an electron. Quantum numbers are a set of 4 numbers.

Name	Symbol	Describes	Found	Possibilities
Principal energy level	n	Main energy level	Shell #2	Whole # > 0
Azimuthal number	l	Subshell shape	s = 0, p = 1, d = 2, f = 3	Whole # < n
Magnetic	m_l	Which orbital	Number line system (middle is 0)	-l → +l
Spin	m_s	Up or down arrow	Up = + ½ Down = - ½	+ or - ½

Chapter Summary

Atoms chemically bond together to form molecules. This tutorial introduces the types of bonds that occur, and theories on how orbitals in the molecule.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Bonding

- Ionic
- Covalent
- Polar Covalent
 - Bond polarity
- Metallic

Lattice energy

Isomers and resonance

Valence Bond theory

Chapter Review

Atoms bond chemically to form molecules.

Types of bonding

There are a number of bond types depending on the elements between them and how the electrons are contributed to the bonding.

Bond type	Happens between	Electrons a
Ionic	Metal & non-metal	Transferred
Covalent	Non-metals	Shared
Polar Covalent	Non-metals	Shared unevenly
Metallic	Metals	pooled

There are general characteristics of each type of bonding:

- **Ionic:** High melting points, most dissolve in water, conduct electricity when dissolved in water, brittle
- **Covalent:** Low melting points, most do not dissolve in water, do not conduct electricity when dissolved in water
- **Polar covalent:** Medium melting points, some dissolve in water, do not conduct electricity when dissolved in water
- **Metallic:** Soft, conduct heat and electricity, do not dissolve in water

Bond Polarity

When nonmetals bond covalently with a large difference in electronegativity, a polar bond is formed. A polar bond has a partial separation of charges. Polar bonds are symbolized with an arrow pointing towards the more electronegative element and a crossed tail by the less electronegative element

Isomers and Resonance

Isomers are compounds with the same chemical formula but different chemical structure. Resonance occurs when a double bond can be placed in more than one place in a structure without creating isomers.

Bonding theory

The **Valence Shell Bonding Theory** assumes that bonds are formed when atomic orbitals overlap. Direct overlap leads to sigma bonds and parallel overlap leads to pi bond. The **Hybridization Theory** says that the sigma bonds are all identical around an atom, therefore the orbitals forming them must be identical. Therefore, any atomic orbitals involved in sigma bonds hybridize into identical orbitals to bond.

Chapter Summary

Lewis structures are a way to represent molecules in two dimensions. Lewis structures show the atoms and their valence electrons and in what ways the electrons are shared or transferred to form bonds. This tutorial shows two methods of drawing Lewis Structures and uses the Valence Shell Electron Pair Repulsion Theory to determine molecular geometry.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Review of bonding

- Covalent versus ionic

Lewis Structures

- Atoms
- Molecules
 - How to arrange the atoms
 - Drawing structures with multiple bonds
 - A second approach to drawing structures
 - Exceptions to the octet rule
 - Ionic compounds
- Formal Charge

Valence Shell Electron Pair Repulsion Theory

- Used to determine electron geometry
- Used to determine molecular geometry

Chapter Review

Atoms bond chemically to form molecules. Lewis structures are a way to represent this bonding on two dimensional paper and determine the molecular geometry of a structure.

Review of bonding

Covalent molecules share electrons while ionic compounds transfer electrons from one atom to another.

Lewis Structures of atoms

The element symbol is drawn to represent the nucleus and core electrons. The valence electrons are drawn around the symbol—one on each side before doubling up.

Exceptions to the Octet Rule

Most atoms are the most stable with 8 electrons in their valence shell, and will bond until this is reached. However, hydrogen and helium can only hold 2 electrons in their valence shell. Boron and Beryllium can be stable with only 6 valence electrons. Any element in the third row or below can hold more than 8 in the empty d subshells.

Arranging atoms in a Lewis Structure

It is often difficult to know in what order to place the atoms. There are some general rules that can be followed:

- For molecules with only 2 elements, arrange the atoms symmetrically
- "COOH" is a carboxylic acid (both O's bond to the C and the H goes on one of the O's)
- Hydrogen and halogens cannot go in the middle
- Write the remaining atoms in the order they appear in the formula
- Write the hydrogen and halogen atoms around the element they are written next to in the formula

Drawing Lewis Structures for covalent compounds

Once the atoms are arranged, a system can be used to complete the Lewis Structure:

- Arrange the atoms as above
- Determine the # of valence electrons for each atom
- Draw the valence electrons—do not double up where a bond is going to form between two atoms
- Count to see if all atoms have full valences
- If two atoms adjacent to each other do not have full valences, move in an electron from each to form a double bond. Repeat for triple bond if necessary.
- If two atoms that are not adjacent to each other need to double bond, try moving a hydrogen to one of them to cause two atoms adjacent to each other to need the double bond.

Another approach to drawing Lewis Structures

There is a second method that is also commonly used to arrive at the same structure:

- Arrange the atoms as above.
- Determine the total # of valence electrons for the whole molecule
- Put one bonding pair between each set of atoms to be bonded.
- Place remaining electrons in lone pairs, starting with the most electronegative element.
- If atoms do not have full valence shells, move a lone pair from an adjacent atom in to double, or triple, bond.

Ionic Structures

Ionic bonds are formed from the transfer of electrons from the metal atom to a non-metal atom or polyatomic ion. When drawing ionic structures, do not draw the atoms as sharing the electrons. Rather, remove the electrons from the

Valence Shell Electron Pair Repulsion Theory

Bonds are made of electrons and electrons are negative and therefore repel each other. Bonds and lone pairs form as far apart from each other as possible. This theory can be used to determine the electron structure (the 3D shape based upon electron regions—bonding regions and lone pair regions—of the central atom) or molecule structure (the 3D shape based on the electron regions, but named after the bonded atoms only).

A = central atom; X = ligands; E = lone pairs

Electron regions	Molecular Formula	Name
2	AX ₂	Linear
3	AX ₃	Trigonal Planar
3	AX ₂ E	Bent
4	AX ₄	Tetrahedron
4	AX ₃ E	Trigonal pyramidal
4	AX ₂ E ₂	Bent
5	AX ₅	Trigonal bipyramidal

5	AX_4E	See-saw
5	AX_3E_2	T-shaped
5	AX_2E_3	Linear
6	AX_6	Octahedron
6	AX_5E	Square pyramidal
6	AX_4E_2	Square planar

Chapter Summary

The periodic table is the main tool chemists use to organize the elements.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Periodic table

- Key sections of the periodic table
- Mnemonic for memorizing 1st 20 elements
- Properties of different key groups

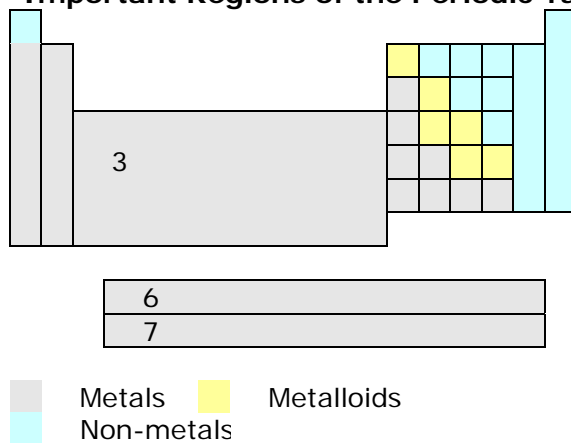
Chapter Review

The periodic table is most likely the key tool for chemists. It organizes the elements, but it also gives a wealth of information.

Key sections of the periodic table

The periodic table is organized in columns, called groups or families, and rows, called periods. There are several groups or periods that have specific names.

Important Regions of the Periodic Table:



1. Alkali Metals
2. Alkaline Earth Metals
3. Transition Metals

4. Halogens
5. Nobel Gases
6. Lanthanides

7. Actinides

8 tall columns = main groups or representative elements

Alkali Metals

They are soft, silver metals with low melting points. They are the most reactive metals, so they are rarely found in elemental form. They have 1 valence electron tends to be lost in bonds to form M^{+1} cations. The metal oxide (M_2O) reacts in water to form bases (MOH). They react with water to form base & H_2 with increasing vigorousness as you move down the group and they form most stable compounds with anions of similar size

Alkaline Earth Metals

They are soft, silver metals with higher melting and boiling points than Alkali metals. They are also rarely found as pure element. Their 2 valence electrons tend to be lost in a bond to form M^{+2} cations.

Metal oxide (MO) react with water to form base ($M(OH)_2$) and the metals react with water & O_2 with increasing ease moving down the group.

Halogens

They have increasing melting & boiling points moving down the group (top 2 are gases, then 1 liquid and 2 solids) and F, Cl, Br & I form diatomic gases (F_2 , Cl_2 , Br_2 , I_2). They form compounds among themselves (interhalogens) such as ClF_3 and BrF_5 which have properties between the two parent halogens. They have 7 valence electrons and tend to gain one more in bonding to form X^{-1} anions.

The non-metal oxides react with water to form acids of increasing strength moving down the group.

Noble Gases

They have very low boiling points and are found as gases. They have full valence shells mean noble gases are unreactive monoatomic gases and they are not inert as once was thought, as xenon does form some compounds (such as XeF_4).

22: The Periodic Table - Properties

Chapter Summary

The periodic table is the main tool chemists use to organize the elements. There are many trends in properties that appear on the periodic table. This tutorial introduces the main sections of the periodic table and explains the trends in properties that occur throughout the table.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Periodicity

- Atomic mass
- Atomic radii
- Electronegativity
- Electron affinity
- Ionization energy

Ionic Radii

Chapter Review

The periodic table is most likely the key tool for chemists. It organizes the elements, but it also gives a wealth of information.

Periodicity

As you move across or down the periodic table, subatomic particles are added. This increases the mass of the elements both across and down the periodic table.

Moving across the periodic table, protons are added to the nucleus while electrons are added to the valence shell. This increase in both the number of positive charges and negative charges increases the attraction between the two. Therefore, when moving across the periodic table, the radius decreases. When moving down the periodic table, protons are again added. But this time, the electrons are added in a completely new valence shell. This new valence shell is shielded from the pull of the protons by all the inner valence shells. Therefore, as you move down the periodic table, atomic radius increases.

Electron affinity (the ease with which an electron is added), electronegativity (pull an electron has on electrons it shares in a bond) and ionization energy (difficulty in removing the outermost electron) are all related to the radius. As radius decreases across the periodic table, all of these properties increase as the electrons are closer to the pull of the

protons. As radius increases down a group, these properties decrease as the electrons are farther from the nucleus.

Ionic Radii

The formation of a cation is due to the loss of electrons. A cation has more protons than electrons and therefore the protons have a large pull on each electron. The radius decreases. An anion is formed from the gain of electrons. Anions have a greater number of electrons than protons. Therefore, the pull of the protons on each electron is lower. The radius of an anion is greater than the parent atom.

23: Nuclear Reactions

Chapter Summary

This tutorial will introduce you to the sub-atomic particles, binding energy of those particles, the nuclear decay processes and how to balance nuclear reactions as well as applications of nuclear reactions.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Sub-atomic particles
Energy that binds particles together
Nuclear decay processes
Balancing nuclear reactions
Applications of nuclear reactions

Chapter Review

Proton:

A positively charged subatomic particle forming part of the nucleus of an atom and determining the atomic number of an element.

Neutron:

A subatomic particle forming part of the nucleus of an atom and having no charge.

Nucleon:

A proton or a neutron.

Atomic Number:

The number, equal to the number of protons in an atom that determines its chemical properties. Symbol: Z .

Atomic Mass:

The mass of an atom expressed in atomic mass units.

Strong (nuclear) Force:

A fundamental force that is associated with the strong bonds between quarks and other subatomic particles.

Weak (nuclear) Force:

One of the four fundamental forces that is associated with nuclear decay.

Binding Energy:

The energy needed to separate the constituent parts of an atom or nucleus.

Mass Defect:

The difference between the mass of an atom and the sum of the masses of its individual components.

Electron-Volt:

A unit of energy.

Mass-Energy Equivalence:

All mass represents an equivalent amount of energy. $1 \text{ amu} = 931 \text{ MeV}$.

Radioactivity:

Emission of radiation as a consequence of a nuclear reaction, or directly from the breakdown of an unstable nucleus.

Half Life:

The time required for half of the nuclei in a sample of a specific isotope to undergo radioactive decay.

Alpha Particle:

A positively charged helium nucleus (consisting of two protons and two neutrons).

Beta Particle:

An energetic electron produced as the result of a nuclear reaction or nuclear decay.

Gamma Particle/Ray:

Very high frequency electromagnetic radiation emitted as a consequence of radioactivity.

Fission:

The process whereby one item splits to become two.

Tomography:

Imaging by sections or sectioning.

24: Chemical Laboratory

Chapter Summary

This tutorial will introduce you to laboratory topics: Lab safety, common lab techniques and data analysis.

Tutorial Features

- Concept map showing inter-connections of concepts.
- Definition slides introduce terms as they are needed.
- Examples given throughout to illustrate how the concepts apply.
- A concise summary is given at the conclusion of the tutorial.

Key Concepts

Lab Safety
Common Lab Techniques
Data Analysis

Chapter Review

Metric System:

System of measurement based on multiples of 10.

SI System:

Systeme International d'Unites (International system of units).

Meniscus:

Curved appearance at the top of liquids in containers due to the attraction between the liquid and the container.

Uncertainty:

The last digit in a measurement is uncertain—each person may see it slightly differently when reading the measurement.

Significant Figures:

Digits that were actually measured and have physical significance. (Also called “significant digits”)