AP Chemistry Rapid Learning Series

Course Study Guide

http://www.RapidLearningCenter.com

AP Chemistry Rapid Learning Series

Introduction

AP Chemistry is a course of fundamentals. Throughout the revised AP curriculum, you will gather many tools and concepts to further your studies in future chemistry. This course studies atoms, molecules and chemical reactions. The series integrates key concepts and problem solving in the areas of AP's six big ideas, which is structured to meet the new AP requiements. The chapter organization and topical flow mirrors the six central big ideas. This rapid course will prepare you for the course itself and the AP exam. The recommended pre-requisites for this course are high school chemistry and high school algebra 1 & 2.

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Tutorial Features

This tutorial series is a carefully selected collection of core concept topics in general chemistry that cover the essential concepts in the course. It features three parts:

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

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.
- 3D visualization of structures.
- Molecular animations of reactions and changes.
- Conceptual explanation of constants.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Problem Solving Drills

- Each tutorial has an interactive problem set covering the material presented in the tutorial. Work out each problem and then check it with the provided answer and complete solution provided at the end.

Condensed Cheat Sheat

- Each tutorial has a one-page cheat sheet that summarizes the key concepts and equations 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 AP Chemistry

Tutorial Summary:

Chemistry is the study of matter and its interactions. AP Chemistry is a two-semester general chemistry at college level taken by high school students who want to gain college credits.

With the revised AP curriculum, the study of chemistry centers on the concepts and practices of the six big ideas.

There are two question types in this AP exam, multiple choice and free-response. Exam strategies on both sections are introduced.

The KUDOS word problem solving method is illustrated to solve chemistry problems.

Helpful tips are given on how to study and be successful in this AP chemistry course.

Tutorial Features:

- Organizational chart showing the six big ideas in AP Chemistry.
- Tips in how to succeed in multiple choice and free-response sections.
- 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.

Concepts Covered:

Definition of Chemistry
What is AP Chemistry?
AP Curriculum – Six big ideas
Multiple choice questions
  - definition
  - tips

Free-response questions
  - definition
  - tips

General problem solving strategy
  - 5-step approach

KUDOS Method
  - K - Known
  - U – Unknown
  - D – Definition (equation)
  - O – Output (calculation)
  - S – Substantiation (double-check and verify)

Tips for studying chemistry

Content Review:

Chemistry is the study of matter and its changes and interactions.

AP Chemistry

- College General Chemistry course offered to High School Students.
- Students can take an exam in May.
- The score on the exam (1-5) and the college you attend determine if you receive college credit for the course.

Multiple Choice Questions

90 minutes - No calculators, only a periodic table and equation sheet, and 60 questions (discrete questions and question sets).

- Move on when you get stuck…you're not expected to know everything!
- Make educated guess – no guessing penalty in AP.
- Get used to working without a calculator.
- Scan all the choices before choosing your answer.
- Try to rephrase things into terms you're more comfortable with.
- Beware of absolutes—there are very few things in chemistry that are absolute!

Free-response Questions

90 minutes total (two parts) - Calculator OK, periodic table & equation sheet. Part A – 3 long questions: Part B – 4 short questions.

- Study the contents on the 5 big ideas.
- Become familiar with the equation sheet throughout the year so that you can quickly find information on test day.
- Show work—the readers can give partial credit for partial correct work.
- Answer the question, only the question.
- Learn to explain the "why".
- Do what you can for partial credit without leaving blank.

General Problem Solving Strategy

Five-Step Process:

- Step 1: Identify what's being given.
- Step 2: Clarify what's being asked. If necessary, rephrase the question.
- Step 3: Select a strategy. Trial & error search, deductive reasoning, knowledge-based, working backwards.
- Step 4: Solve using the strategy.
- Step 5: Review the answer.

KUDOS Method

K (Known):
- Use units to identify information.
- Write information symbolically.
- Look for implied information.
- Write out chemical equations.

U (Unknown)
- What is the problem looking for?
- Write information symbolically.

D (Definition)
- Find equalities to convert.
- Choose & Re-arrange equations.
- Look for missing information in other places.
- If you cannot find enough information, re-evaluate your plan.

O (Output)
- Plug in values to the equations (use constants as needed).
- Check unit cancellation & perform the calculation.

S (Substantiation)
- Check validity of your answer.
- Check units.
- Check significant figures.

02: Basic Skills – Units and Measurements

Tutorial Summary:

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.

Chemistry is an experimental science; therefore it is necessary to take careful measurements. Measurements should always include one more decimal place than the instrument indicates for certain—this last decimal place should be a "0" if the measurement is "on the line" and a "5" if the measurement is "in-between the lines." Most measurements are understood to be +/- 1 in the last decimal place, unless another uncertainty value is given with the measurement.

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.

There are many other "Basic skills" needed in chemistry, but they will be introduced throughout the tutorial series when they are needed.

Tutorial Features:

- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

The Metric and SI Systems
  - Metric prefixes
  - SI units
  - Common non-SI units

Measurement and Uncertainty
  - Tools common in Chemistry Labs
  - Taking measurements
  - Uncertainty

Significant Figures
  - Why they're used
  - How to count them

Fundamental Constants used in Chemistry

Brief overview of topics to be included in future tutorials:
  - The language of chemistry
  - The periodic table
  - Dimensional analysis
  - Scientific notation
  - Use of scientific calculators

Content Review:

Chemistry is an experimental science, therefore it is necessary to be able to work with units and measurements accurately.

Metric System

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

Metric Prefixes commonly used in chemistry

Prefix	Symbol	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

Taking Measurements

Measurements must be taken accurately. Always write down one more decimal place than the instrument tells for certain—a “0” if it’s “one the line” and a “5” if it’s “between the lines.”

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

03: Math Review for AP Chemistry

Tutorial Summary:

Math skills are crucial throughout chemistry. This tutorial reviews basic algebra needed in chemistry calculations. Writing answers with the correct number of significant figures is taught, along with writing and reading numbers in scientific notation. Performing calculations with exponents, including those in scientific notation, is illustrated. Logarithms, both base 10 and natural logs, is taught. The quadratic equation is demonstrated. Finally, calculator survival tips are given to ensure the answer you type in what you intend. All mathematical concepts are illustrated with chemistry applications—each one is actually used in chemistry calculations.

Tutorial Features:

- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

Algebra
  - Solving for a variable in addition/subtraction and multiplication/division

Calculating with significant figures
  - Addition/subtraction
  - Multiplication/division

Scientific Notation
  - Writing in scientific notation
  - Reading scientific notation

Exponents rules
  - Calculations with exponents
  - Calculations with scientific notation

Logarithms
  - Logarithms with base 10
  - Natural logarithms

Quadratic Equation
Calculator tips

Content Review:

Math skills are needed throughout a chemistry course.

Algebra

Algebra is used to solve equations by un-doing whatever is being done to an unknown variable. For example, if an equation has “x+2” then you would subtract “2” to solve for “x”. Everything that is done to one side must be done to the other side of the equation as well.

Calculations with Significant Figures

You cannot become more precise after completing calculations than the original data was. Therefore, it is important to write the answer with the correct number of significant figures. When adding and subtracting with significant figures, you write the answer with the least number of decimal places that are in the problem. When multiplying and dividing, write the answer with the least number of significant figures as is in the problems.

Scientific Notation

Scientific notation is a way of writing large or small numbers as a multiple of 10. The decimal place is always placed behind the first non-zero number and the number of times the decimal point was moved to get there is used as the exponent of 10. Positive exponents represent large numbers (>1) and negative exponents represent small numbers (<1).

Exponents

Several rules are used for exponents:

- Anything to the 1st power = the original number
- Anything to the power of 0 = 1
- When multiplying numbers with exponents, if the #'s are the same, you can add the exponents and put the total with the original number
- When dividing numbers with exponents, if the #'s are the same, you can subtract the exponents and put the difference with the original number
- When taking a power to a power, multiply the powers
- A negative power indicates that it's on the opposite side of the fraction

There are similar rules when working with scientific notation numbers:

- Addition with same powers of 10: Add the numbers and keep the power of 10 the same.
- Subtraction with the same powers of 10: Subtract the numbers and keep the power of 10 the same.
- Multiplication: Multiply the numbers and add the powers of 10
- Division: Divide the numbers and subtract the powers of 10
- Power: Take the number to that power and multiply the power of 10 by the power
- Roots: Take the root of the number and divide the power of 10 by the root

Logarithms

Logarithms are a way of counting in multiples of a base number. If $x = \log_b y$ then $y = b^x$. If no base is specified, it's assumed to be 10. A natural log (ln) uses the base of "e" (2.313).

Quadratic Formula

The quadratic formula is used to solve for a variable when there is the variable to the 2nd power and to the 1st power in the same equation. If $ax^2 + bx + c = 0$ then $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

Calculator Tips

People often get incorrect answers simply from a mistake in the way they enter numbers into their calculator. When dividing by more than one number, use the ÷ button each time. When entering scientific notation, always use the EE (or EXP) button rather than entering (^10). Be sure to use parenthesis around addition and subtraction when combining with multiplication and division, and also when taking a value (especially a negative value) to a power.

04: Dimensional Analysis

Tutorial Summary:

Dimensional analysis is a technique for converting units based upon equalities. The technique is the basis for stoichiometry—a key chemistry calculation—presented later in the series.

Tutorial Features:

- Problem-solving techniques are used to work out the example problems.
- 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 concept
- Animated examples—worked out step by ste
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

What is dimensional analysis?
  - What is the principle behind dimensional analysis?

What equalities are used commonly in chemistry?
How is dimensional analysis performed?
How is dimensional analysis used in chemistry?

Content Review:

Dimensional analysis is used to convert units.

The Principle Behind Dimensional Analysis

The guiding principle of dimensional analysis is that you can multiply anything by “1” without changing the meaning. An equality set into a fraction formation = 1. For example, if $x = y$, then $x/y = 1$ and $y/x = 1$. Therefore, the equalities can be set into fractions and multiplied to convert units.

Another concept necessary to understanding dimensional analysis is that units that are on the top and bottom of an expression cancel out.

Equalities Commonly used in Dimensional Analysis

Several equalities are used often in chemistry.

Equalities commonly used in chemistry	
4.18 J	1.00 cal
1 Å	10^{-10} m
1 cm^3	1 mL
1 dm^3	1 L
1 in	2.54 cm
1 kg	2.2 lb
1 atm	101.3 kPa
1 atm	760 mm Hg
1 mole	6.02×10^{23} pieces

Metric prefixes are also used to form equalities between different metric units.

Dimensional Analysis

To work dimensional analysis problems:

- Write your known down on the left side
- Write down “=___________ [desired unit]” at the right side

- Identify equalities that will get you from the known information to the desired unit. If there is no equality that involves both the known and unknown, you'll have to find more than one to more than one step.
- Arrange the equalities into a fractional form so that the known unit will cancel out and the desired unit will be left.
- Multiply across the top of the expression and divide numbers on the bottom.

05: Matter and Energy

Tutorial 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.

Common misconceptions on the process of science are explained correctly: (1) The Scientific Law, (2) Theory versus Law and (3) Prediction versus Hypothesis.

Helpful tips are given on how to study and be successful in a chemistry course.

Tutorial Features:

- Concept map showing inter-connections of concepts introduced.
- 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.
- Organizational chart showing classifications of matter.
- Molecular visualizations of different types of matter.
- Common misconceptions pointed out for the student to be aware of.

Concepts Covered

- 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 Properties
    - Macroscopic versus Microscopic Properties
- Scientific Process
    - “The Scientific Method” versus Scientific Processes
    - Theory versus Law
    - Prediction versus Hypothesis

Content 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.

Scientific Processes

There are many paths to follow when undertaking “science”—there is no one scientific method. Science involves observing, posing questions, forming possible explanations (hypothesis), experimenting, processing/analyzing data, looking for trends, more formation of possible explanations or question posing. Scientific processes form theories (which attempt to explain observed behavior) and laws (which describe or predict behavior, and are usually mathematical). A theory cannot become a law—one explains why and one describes what. A hypothesis is a proposed explanation for why something will occur (that may become a theory with enough evidence), while a prediction is simply a guess at what will happen—it does not attempt to say “why” it will happen.

06: Atoms and Molecules

Tutorial 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 new concepts in this tutorial and those previously introduced.
- Definition slides introduce terms as they are needed.
- Visual representation of concepts.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

Atoms
Ions
Element Symbols
Isotopes
Atoms, Elements, and Molecules

Content 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 is 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.

07: Writing Formulas

Tutorial 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:

- Nomenclature is broken down into separate types
- Organizational chart at end to guide type selection
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

Writing chemical formulas

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

Content Review:

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

Type 1: Binary Ionic Compounds
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 Compounds
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 Compounds
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.

08: Naming Chemicals

Tutorial 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:

- Nomenclature is broken down into separate types
- Organizational chart at end to guide type selection
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

Naming chemical formulas

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

Content Review:

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

Type 1: Binary Ionic Compounds

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 Compounds

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 are 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 Compounds

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.

09: The Mole

Tutorial 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 new concepts in this tutorial and those previously introduced.
- Definition slides introduce terms as they are needed.
- Visual representation of concepts.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

The Mole
  - Mole relationships

Molar Mass
  - Calculating
  - Using in mole conversions

Percent Composition
Empirical formulas
Molecular formulas

Content 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 part / whole × 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.

10: Atomic Structure and Electron Configuration

Tutorial 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:

- The periodic table is used as visual tool to remember orbital orders
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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):

1s

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

11: The Periodic Table and Periodicity

Tutorial 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:

- The trends aren't just given…they are explained in terms of one another and linked together
- Mnemonic is given for easy memorizing of 1st 20 elements
- Animations of the periodic table
- Visual representation of trends of the periodic table
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

Periodic table
  - Key sections of the periodic table
  - Mnemonic for memorizing 1st 20 elements

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

Ionic Radii

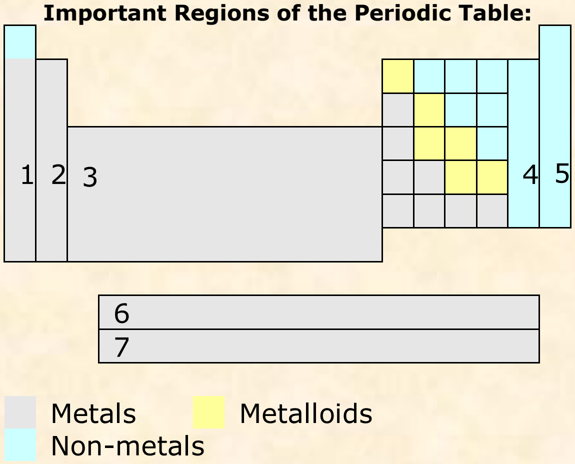
Content Review:

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. Actidines

8 tall columns = main groups or representative elements

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.

12: Chemical Bonding and Molecular Structure

Tutorial 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:

- Molecular animations of orbitals coming together to bond
- 3D visualization of hybrid and molecular orbitals
- Animation of electrons moving from atomic orbitals to molecular orbitals
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Isomers and resonance

Bonding theories
  - Valence shell bonding
  - Hybridization
  - Molecular orbital
    - Bonding and anti-bonding orbitals
    - Bond order

Content Review:

Atoms bond chemically to form molecules.

Types of Bonding

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.

Bond type	**Happens between**	**Electrons are**
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. The **Molecular Orbital Theory** says that rather than overlapping atomic orbitals, or hybridized atomic orbitals, new orbitals formed in the molecule are separated. Some of the new orbitals formed pull the two nuclei together and are bonding orbitals. However, some of them pull the two orbitals apart and are anti-bonding orbitals. The bond order is the number of bonding orbital electrons – the number of antibonding orbital electrons divided by 2. The bond order can be used to determine the type of bond (single, double, triple).

13: Lewis Structure

Tutorial 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:

- Electrons are animated as they move to be shared or transferred
- 3D visualization of molecular geometry
- Two approaches to drawing Lewis Structures are introduced
- Animation of electrons moving from atomic orbitals to molecular orbitals
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Valence Shell Electron Pair Repulsion Theory
  - Used to determine electron geometry
  - Used to determine molecular geometry

Content 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 valance 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 metal atom and add the electrons to the non-metal atom.

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_2E	Bent
4	AX_4	Tetrahedron
4	AX_3E	Trigonal pyramidal
4	AX_2E_2	Bent
5	AX_5	Trigonal bipyramidal
5	AX_4E	See-saw
5	$AX_3 E_2$	T-shaped
5	AX_2E_3	Linear
6	AX_6	Octahedron
6	AX_5E	Square pyramidal
6	$AX_4 E_2$	Square planar

14: Gas Laws

Tutorial 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:

- Animated visualization of gas particles and gas behavior
- Conceptual explanations of gas behavior along with mathematical gas laws
- Problem solving strategy for solving gas law problems
- Combination several gas laws into 1 for less confusion in choosing the appropriate law
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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: $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$ 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: $\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT$

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

15: Liquids and Solids

Tutorial 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:

- 3D visualizations of the solid state structures
- Animations of intermolecular forces (such as the formation of a temporary dipole in London Dispersion Forces)
- Animated visualization of Vapor Pressure
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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.

16: Solutions

Tutorial Summary:

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

Tutorial Features:

- Animations of concepts such as factors affecting solubility, dilutions, vapor pressure of a solid and colloids exhibiting the Tyndall Effect
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

Process of solution formation
Factors affecting solubility
Concentration calculations
  - Expressing concentration
  - Dilution calculations
  - Using concentration in stoichiometry

Electrolyte solutions
Colligative properties
  - Vapor pressure
  - Boiling point
  - Freezing point
  - Effect of electrolytes on colligative properties

Colloids
  - Tyndall Effect

Content 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: $\% \, mass = \frac{mass \; solute}{mass \; solution} \times 100$

The mass units must match!

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

The volume units must match!

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

The volume unit is mL

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

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

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!

Colligative Properties

A colligative property is a property that depends on the number of solute particles in the sample. The vapor pressure of a solution is lower than the pure solvent because the number of solvent particles on the top layer that can evaporate is lower. Because the vapor pressure is lower, the boiling point of a solution is always the higher than the pure solvent and the freezing point is always lower than the pure solvent. An electrolyte solution, one in which the solute breaks apart into multiple ions which allow electricity to be conducted, has an even greater change in vapor pressure, boiling point or freezing point because there are more particles in the solution than molecules added to the solution.

Colloids

Colloids are solutions with solute particles large enough to scatter light. They exhibit the Tyndall Effect, where light is seen traveling through and spreading out as it travels through colloid.

17: Chemical Reactions

Tutorial 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:

- Molecular animations of chemical reactions
- Animations demonstrating how to determine products of a chemical reaction
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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 and 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).

18: Balancing Equations

Tutorial 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:

- Instructions on how to choose the atom to begin with is given
- Multiple methods of balancing equations are worked out in animation.
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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 equals 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 H_2O 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 a 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.

19: Stoichiometry

Tutorial 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:

- Use of real-life analogies to explain stoichiometric concepts
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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.

20: Kinetics

Tutorial 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:

- Animation of conditions necessary for a reaction to occur
- Simultaneous animation of a reaction and a reaction coordinate diagram
- Molecular animations of elementary steps in a reaction mechanism
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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 steps show 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 law 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.

21: Equilibrium

Tutorial 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:

- Molecular animation of establishment of dynamic equilibrium
- Animation of equations to show a change in the ratio of product to reactants
- Use of "ICE Chart" problem solving technique for solving equilibrium problems
- Application of Le Chatelier's principle to the manufacturing industry
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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.

22: Acids and Bases

Tutorial 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:

- Molecular animations of acid-base reactions
- Use of ICE equilibrium technique to solve weak acid problems
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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
Titrations

Content 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, HNO_3, $HClO_3$, $HClO_4$. 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 it's "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 × 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. $pH = -\log[H_3O^{+1}]$

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.

23: Thermodynamics

Tutorial 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 new concepts in this tutorial and those previously introduced.
- Definition slides introduce terms as they are needed.
- Visual representation of concepts.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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:

$$\Delta Hrxn = \sum H_f\, prod - \sum H_f\, react$$

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

24: Electrochemistry

Tutorial 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:

- Review of redox content from earlier tutorials
- Particle animation of voltaic cell and explanation of components of cell.
- 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.
- Animated examples—worked out step by step.
- A concise summary is given at the conclusion of the tutorial.

Concepts Covered:

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

Content 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 H_2O 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 a 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 reaction, 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* = cathode – 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: $\Delta G = -n \times F \times EMF$ ΔG = free energy (in J); n = # of moles of electrons transferred; F = 1 Faraday; EMF = 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 J/mole×K; T = temperature (in Kelvin); n = moles electrons transferred; F = 1 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.