High School Chemistry Rapid Learning Series

Course Study Guide

Introduction to High School Chemistry

High School Chemistry is often a student’s first exposure to chemistry. You may not even be sure what “chemistry” really is. Many High Schools and Colleges are now requiring students to take High School Chemistry. This series will introduce you to the basic concepts and problem solving included every High School Chemistry Course, typically a two-semester class.

Learning chemistry is about the understanding of the key concepts and the application of these concepts onto problem solving in chemistry. These two key ingredients are also the focus of this course. Unlike traditional textbooks, this Rapid Learning series will introduce you the chemistry via visual learning and smart teaching.

There are total of 24 chapters in this course, each chapter will require the study of three components: core tutorial, problem drill and review sheet.

Let us put some real effort to learn chemistry, one hour at a time. Execute your study plan and carry your learning to the finish line.

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Tutorial Series Features

This tutorial series is a carefully selected collection of core concept topics in high school chemistry that covers the essential concepts and problem solving. It features three core parts:

- Concept Tutorials – 24 essential topics
- Problem-Solving Drills – 24 practice sets
- Super-Review Cheat Sheets – 24 super review sheets

Core Component#1: Visual Tutorials (x24)

- Self-contained tutorials, not just an outline
- Concept map showing inter-connections of related concepts
- Definition slides introduced contextually
- Visual representation of new concepts
- Molecular visualization of structures
- Molecular animations of reactions
- Animated problem solving step by step
- A concise summary

Core Component #2: Problem-Solving Drills (x24)

- Chapter-related problems
- Feedback-based drills
- Intereactive on every click
- Complete solutions at the end
- Review and retake at will

Core Component #3: Super-Review Sheets (x24)

- Key concepts and formulas for the chapter
- One-page summary sheet per chapter
- Printable in PDF

Supplemental Component #1: Printable eBooks (x24)

- Well formatted printable tutorials
- One eBook per chapter with total of 24
- Easy reading for re-cap and final review

Supplemental Component #2: Audiobooks (x24)

- Learning on-to-go
- 24 mp3 audiobooks for 24 chapters
- Enhancement to visual learning
- Immersion study anywhere anytime

01: Introduction to Chemistry

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
- Tips for Studying Chemistry

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

02: Math for 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, are 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.
- Examples of the math techniques use chemistry applications.

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
    - Calculations with scientific notation
- Logarithms
    - Logarithms with base 10
- 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).

There are 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.

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 you enter a 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.

03: 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

- 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.
- Problem-solving techniques are used to work out the example problems.

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.

04: Problem Solving in Chemistry

Tutorial Summary

You will encounter problems throughout your chemistry course. A "problem" is anytime you don't know how to get from the known information to the unknown information. Having a systematic method to "attack" problems can make them much easier. A general problem-solving technique is introduced in this tutorial as well as a specific technique for word problems. Tips for preparing for and taking tests are also given.

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.
- The problem-solving technique is applied to chemistry applications.

Concepts Covered

- General problem solving technique
- Problem solving technique for word problems
- Preparing for tests
- Taking tests

Content Review

Problem-solving can be a very difficult task. But when a specific "attack" is used, the problems can be solved much more easily.

General Problem-Solving Technique

A technique for solving any type of problem involves:

- Identify what's being given.
- Clarify what's being asked.
- Select a strategy (trial and error, search, recall, deductive reasoning, working backwards, etc).
- Solve using the selected strategy.
- Review your answer.

Problem-Solving Technique for Word Problems

A problem-solving strategy that turns the above "general" strategy into a strategy specifically for solving word problems is the KUDOS method:

- K = Known (identify the known).
- U = Unknown (identify the unknown).
- D = Definition (identify definitions, equations, equalities, constants needed).
- O = Output (use your definitions to output an answer).
- S = Substantiation (check appropriateness, units and significant figures of your answer).

Preparing for a Test

There are several tips to help you prepare for a test:

- Stay ahead of the game. Don't wait until the last minute. If you wait until the last minute to prepare, there won't be time to seek out help if you encounter problems.
- Make a cheat sheet. Even if you can't use it on the exam, making a concise "cheat sheet" helps you focus in on key information and is a great study technique.
- Know the format of the exam and what information is "test-able" (what you'll be responsible for). Ask your instructor well in advance for this information.
- Make a mock-exam. Select the key information and make a practice exam for yourself.

- Attend any review sessions offered, even if they are optional. Often, key information is given in these sessions that are not said in the general class period.
- Get help early. If you have any problems, seek out your instructor ahead of time. Instructors do not like last minute cramming questions.

Taking a Test

Several tips can help improve your test scores:

- Arrive early and be prepared. Rushing can increase anxiety and decrease performance
- Listen and read instructions carefully. Your instructor may give changes to the questions or instructions. Don't assume you know what a problem is asking…always read the instructions!
- Memory dump. On the back or top of the test, write all information you've been "storing" in memory. This allows a "bank" for you to refer to as you work through the test and prevents anxiety about forgetting information.
- Skim the test and make a plan. Budget your time accordingly.
- Answer questions sequentially. Often information from previous problems is used later in the test—try not to skip around.
- Apply the guessing rule. Don't randomly guess, but if you can eliminate a few of the selections, guess among the remaining.

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: Pure Substances—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 Chemical 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

- 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.
- Nomenclature is broken down into separate types.
- Organizational chart at end to guide type selection.

Concepts Covered

- Writing chemical formulas
    - Binary Ionic Formulas
    - Polyatomic Ionic Formulas
    - Ionic with Multivalent Metals Formulas
    - Binary Covalent Formulas
    - Acids Formulas

Content 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
Binary covalent compounds are 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 Chemical Compounds

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

- 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.
- Nomenclature is broken down into separate types.
- Organizational chart at end to guide type selection.

Concepts Covered

- Naming chemical formulas
    - Binary Ionic Naming
    - Polyatomic Ionic Naming
    - Ionic with Multivalent Metals Naming
    - Binary Covalent Naming
    - Acids Naming

Content Review

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

Type 1: Binary Ionic
Binary ionic compounds are 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
Binary covalent compounds are 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: Counting Molecules — 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: 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

- 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.
- Molecular animations of chemical reactions.
- Animations demonstrating how to determine products of a chemical reaction.

Concepts Covered

- Components of a chemical reaction
- Common types of chemical reactions
    - Composition
    - Decomposition
    - Single replacement
    - Double replacement
    - Neutralization
    - Precipitation
    - Redox
- Determine products of a double replacement reaction
- Using solubility rules to determine a precipitate

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.

Double replacement reactions and precipitations

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

11: 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.

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.
- Instructions on how to choose the atom to begin with is given.
- Multiple methods of balancing equations are worked out in animation.

Concepts Covered

- Using The Law of Conservation of Mass to balance the equations
- Choosing which atom to begin balancing with
- Inspection 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.

12: The Math of Chemical Reactions - 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

- 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.
- Use of real-life analogies to explain stoichiometric concepts.

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.

13: 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

- 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.
- The periodic table is used as visual tool to remember orbital orders.

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

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}

14: The Periodic Table

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

- 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.
- 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.

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

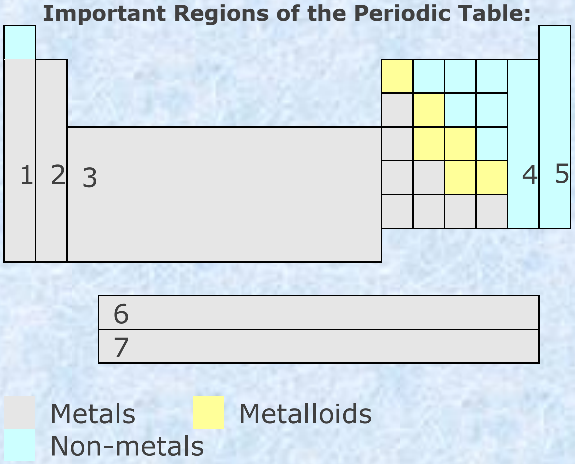
Content Review

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.

15: Chemical Bonding

Tutorial Summary

Atoms chemically bond together to form molecules. This tutorial introduces the types of bonds that occur, and the most simple bonding theory—Valence Bonding Theory.

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.
- Molecular animations of orbitals coming together to bond.

Concepts Covered

- Bonding
  - Ionic
  - Covalent
  - Polar Covalent
    - Bond polarity
  - Metallic
- Valence Shell Bonding Theory

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

Valence Shell 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.

16: Drawing Molecules

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.

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.
- Electrons are animated as they move to be shared or transferred.
- Two approaches to drawing Lewis Structures are introduced.

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

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 you draw ionic structures, do not draw the atoms as sharing the electrons. Rather, remove the electrons from the atom (cation) to another atom (anion). Place the (+) superscript on the cationic atom and (-) superscript on the anionic atom, i.e. Na^+Cl^-.

17: Molecular Geometry

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 uses the Valence Shell Electron Pair Repulsion Theory to determine molecular geometry.

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.
- 3D visualization of molecular geometry.

Concepts Covered

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

Content Review

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

18: 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 and several gas laws.

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.
- 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.

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

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.

19: 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 and colloids. It also reviews using concentration in stoichiometric 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.
- Animations of concepts such as factors affecting solubility, dilutions and colloids exhibiting the Tyndall Effect.

Concepts Covered

- Process of solution formation
- Factors affecting solubility
- Concentration calculations
    - Expressing concentration
    - Dilution calculations
    - Using concentration in stoichiometry
- Electrolyte solutions
- 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. The solute and solvent then 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!

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.

20: Reaction Rates and Equilibrium

Tutorial Summary

Kinetics is the study of reaction rates. This tutorial introduces factors affecting the rate of reaction. 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. Le Chatelier's principle will be introduced.

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.
- 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.
- Molecular animation of establishment of dynamic equilibrium.
- Animation of equations to show a change in the ratio of product to reactants.

Concepts Covered

- Collision Theory
- Reaction Coordinate Diagrams
- Factors affecting rate
- 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
- Le Chatelier's principle

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.

Equilibrium

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.

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.

21: 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, along with the acid or base properties of the salts produced. The pH scale and pH calculations will be introduced, along with buffers and titrations.

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.
- Molecular animations of acid-base reactions.

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
- 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^{+1})
- **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.

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.

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.

22: Energy & Physical Changes

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.

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

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.

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.

23: Energy & Chemistry Changes

Tutorial Summary

Thermodynamics is the study of heat changes during processes. 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 chemical changes
    - Heat of formation
    - Hess's Law
    - Heat of reaction
    - Calorimetry
- Entropy
- Free Energy

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.

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 explains 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.

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.
- Review of redox content from earlier tutorials.
- Particle animation of voltaic cell and explanation of components of cell.

Concepts Covered

- Oxidation states
- Redox reaction
- Balancing redox reactions
    - Oxidation Number method
    - Half-Reaction method
- 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.