Class Notes for Chemistry 1

(including Pre-AP[®] Chemistry)

in Plain English

Mr. Bigler Lynn Public Schools April 2025



http://www.mrbigler.com/Chemistry-1/Notes-Chemistry-1.pdf

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ISBN-13: 978-1517557300 ISBN-10: 1517557305 This is a set of class notes for a first-year high school chemistry course. These notes can be used for any honors or CP1 chemistry course by omitting information that is specific to the higher-level course.

This hardcopy is provided so that you can fully participate in class discussions without having to worry about writing everything down.

While a significant amount of detail is included in these notes, they are intended to supplement the textbook, classroom discussions, experiments and activities. These class notes and the textbook discussion of the same topics are intended to be complementary. In some cases, the notes and the textbook differ in method or presentation, but the chemistry is the same. There may be errors and/or omissions in any textbook. There are almost certainly errors and omissions in these notes, despite my best efforts to make them clear, correct, and complete.

About the Homework Problems

The homework problems include a mixture of easy and challenging problems. (Remember that these notes are intended for use in both regular-level and honors-level chemistry classes.) *The process of making yourself smarter involves challenging yourself, even if you are not sure how to proceed.* By spending at least 10 minutes attempting each problem, you build neural connections between what you have learned and what you are trying to do. Even if you are not able to get the answer, when we go over those problems in class, you will reinforce the neural connections that led in the correct direction.

Answers to the problems are often provided so you can check your work and see if you are on the right track. Do not simply write those answers down, in order to receive credit for work you did not do. This will give you a false sense of confidence, and will actively prevent you from using the problems to make yourself smarter. *You have been warned.*

Using These Notes

As we discuss topics in class, you will want to add your own notes to these. If you have purchased this copy, you are encouraged to write directly in it, just as you would write in your own notebook. If this copy was issued to you by the school and you intend to return it at the end of the year, you will need to write your supplemental notes on separate paper. If you do this, be sure to write down page numbers in your notes, to make cross-referencing easier.

You should bring these notes to class every day, because lectures and discussions will follow these notes, which will also be projected onto the SMART board.

Chemistry 1

Note to Teachers

These notes, and the course they accompany, are designed to follow both the 2016 Massachusetts Curriculum Frameworks, and the 2016 Massachusetts Curriculum Frameworks, which are based on the Next Generation Science Standards (NGSS). The notes also utilize strategies from the following popular teaching methods:

- Each topic includes Mastery Objectives and Success Criteria, as presented in the *Studying Skillful Teaching* course, from Research for Better Teaching (RBT).
- Each topic includes Tier 2 vocabulary words and language objectives, from the Rethinking Equity and Teaching for English Language Learners (RETELL) course.
- Notes are organized in two-column notes format, with a top-down web at the beginning of each unit and a page for students to summarize the unit at the end, from Keys To Literacy.

The order of topics in a chemistry course is a hotly-debated subject. These notes and the course they accompany are organized as follows:



The most controversial decision was to place gases & gas laws at the beginning of the course, right after matter. The rationale is to keep the macroscopic study of matter together, and to have a topic that is rich in lab experiments and demonstrations early in the course. One downside to this approach is that moles need to be introduced in a rudimentary fashion early in the course, and re-introduced and expanded on later.

These notes would not have been possible without the assistance of many other people. I cannot mention everyone, but I would particularly like to thank:

- Every student I have ever taught, for helping me learn how to teach, and how to explain and convey challenging concepts.
- Every teacher I have worked with, for their kind words, sympathetic listening, helpful advice and suggestions, and other contributions great and small that have helped me to enjoy and become competent at the profession of teaching.
- Many of the department heads and principals I have worked with, for mentoring and encouraging me and allowing me to develop my own teaching style. In particular, Mark Greenman, Marilyn Hurwitz, Scott Gordon, Barbara Osterfield, Wendell Cerne, John Graceffa, Maura Walsh, Lauren Mezzetti, and Tom Strangie.
- The chemistry teachers I have worked with over the years, who have generously shared their time, expertise, and materials. In particular, Kathy McGrane, who shared her teaching notes with me, which ultimately inspired me to create these; Kirstin Bunker, who gave significant help with the Keys to Literacy® top-down webs; Beth Hocking, with whom I have had more discussions than I can count about creative ideas for teaching chemistry; and Harvey Gendreau, whose contributions to the AP Chemistry email discussion forums and ChemEd conferences have taught me much about the chemistry that high school students need to understand in order to be successful.
- Everyone else who has shared insights, stories, and experiences in chemistry, many of which are reflected in some way in these notes.

I am reminded of Sir Isaac Newton's famous quote, "If I have seen further it is by standing on the shoulders of giants."

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MA Curriculum Frameworks for Chemistry

Except where denoted with (MA), these standards are the same as the Next Generation Science (NGSS) Standards

Standard	Topics	Chapter(s)
HS-PS1-1	Periodic table & periodic trends: ionization energy, atomic & ionic radius	9
HS-PS1-2	Types of chemical reactions. Predicting products. Intermolecular Forces (IMF) vs. physical state at room temperature	12, 14, 15
HS-PS1-3	IMF vs. bulk properties (melting point/boiling point, density, vapor pressure, etc.)	11, 12
HS-PS1-4	Energy of reaction, heat of formation	18
HS-PS1-5	Reaction kinetics as related to Kinetic Molecular Theory (KMT) & collision theory	19
HS-PS1-6	Equilibrium, Le Châtelier's principle, as related to KMT	19
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HS-PS1-10(MA)	Oxidation/reduction (REDOX) reactions, oxidation numbers, predicting products of REDOX reactions, conceptual electrochemistry	16
HS-PS1-11(MA)	Mixture separation based on chemical & physical properties (<i>e.g.,</i> chromatography, distillation, centrifuging, precipitation reactions)	2, 4, 14
HS-PS2-6	Molecular structures of ionic compounds, acids, bases, metals & polymers	10, 11
HS-PS2-7(MA)	Solvent polarity & why ions dissolve in polar solvents	11, 12, 14
HS-PS2-8(MA)	KMT & gases (electrostatic forces, interactions between molecules in solids, liquids & gases), combined gas law	5
HS-PS3-4b	Conservation of energy with respect to enthalpy, entropy, and free energy (conceptual)	18

MA Science Practices

Practice	Description
SP1	Asking questions.
SP2	Developing & using models.
SP3	Planning & carrying out investigations.
SP4	Analyzing & interpreting data.
SP5	Using mathematics & computational thinking.
SP6	Constructing explanations.
SP7	Engaging in argument from evidence.
SP8	Obtaining, evaluating and communicating information.

Two-Column Notes

Unit: Introduction

Details

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Use the two-column note-taking system to take effective notes, or add to existing notes.

Success Criteria:

• Notes are in two columns with appropriate main ideas on the left and details on the right.

Language Objectives:

• Understand and describe how two-column notes are different from other forms of note-taking.

Notes:

The two-column note-taking system is based on the Cornell note-taking system, which was developed in the 1950s at Cornell University. Besides being a useful system for note-taking in general, it is an especially useful system for interacting with someone else's notes (such as these) in order to get more out of them.

The main features of the two-column note-taking system are:

- 1. The main section of the page is for the details of what actually gets covered in class.
- 2. The left section (officially 2½ inches, though I have shrunk it slightly to 2¼ inches for these notes) is for "big ideas"—the organizational headings that help you organize these notes and find details that you are looking for. These have been left blank for you to add throughout the year, because the process of deciding what is important is a key element of understanding and remembering.
- 3. In the Cornell Notes system, the bottom section (2 inches) is officially for you to add a 1–2 sentence summary of the page in your own words. This is always a good idea, but you may also choose to use that space for other things you want to remember that aren't in these notes.

Big Ideas	Details Unit: Introduction
	How to Get Nothing Worthwhile Out Of These Notes
	Because this book serves as a combination of your textbook and a set of notes, you may be tempted to sleep through class because "it's all in the book," and then use these notes look up how to do the homework problems when you get confused. If you do this, you will learn very little chemistry, and you will find this class to be both frustrating and boring.
	How to Get the Most Out Of These Notes
	These notes are provided so you can pay attention and participate in class without having to worry about writing everything down. However, because active listening, participation and note-taking improve your ability to understand and remember, it is important that you interact with these notes and the discussion.
	The "Big Ideas" column on the left of each page has been deliberately left blank. This is to give you the opportunity to go through your notes and categorize each section according to the big ideas it contains. Doing this throughout the year will help you keep the information organized in your brain—it's a lot easier to remember things when your brain has a place to put them!
	If we discuss something in class that you want to remember, mark or highlight it in the notes! If we discuss an alternative way to think about something that works well for you, write it in! You paid for these notes—don't be afraid to use them!
	There is a summary section at the bottom of each page, and a Keys to Literacy [®] style summary template at the end of each chapter. Utilize both of these. If you can summarize something, you understand it; if you understand something, it is much easier to remember.

Unit: Introduction

Details

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Use information from the organization of a textbook to take well-organized notes.

Success Criteria:

- Section headings from text are represented as main ideas.
- All information in section summary is represented in notes.
- Notes include page numbers.

Language Objectives:

• Understand and be able to describe the strategies presented in this section.

Notes:

If you read a textbook the way you would read a novel, you probably won't remember much of what you read. Before you can understand anything, your brain needs enough context to know how to file the information. This is what Albert Einstein was talking about when he said, "It is the theory which decides what we are able to observe."

When you read a section of a textbook, you need to create some context in your brain, and then add a few observations to solidify the context before reading in detail.

René Descartes described this process in 1644 in the preface to his *Principles of Philosophy*:

"I should also have added a word of advice regarding the manner of reading this work, which is, that I should wish the reader at first go over the whole of it, as he would a romance, without greatly straining his attention, or tarrying at the difficulties he may perhaps meet with, and that afterwards, if they seem to him to merit a more careful examination, and he feels a desire to know their causes, he may read it a second time, in order to observe the connection of my reasonings; but that he must not then give it up in despair, although he may not everywhere sufficiently discover the connection of the proof, or understand all the reasonings—it being only necessary to mark with a pen the places where the difficulties occur, and continue reading without interruption to the end; then, if he does not grudge to take up the book a third time, I am confident that he will find in a fresh perusal the solution of most of the difficulties he will have marked before; and that, if any remain, their solution will in the end be found in another reading."

Reading & Taking Notes from a Textbook Page: 12

Big Ideas	Details Unit: Introduction
	The following 4-step system takes about the same amount of time you're used to spending on reading and taking notes, but it will probably make a tremendous difference in how much you understand and remember.
	 Make a two-column notes template. Copy the title/heading of each section as a big idea in the left column. (If the author has taken the trouble to organize the textbook, you should take advantage of it!) Leave about ¼ to ½ page of space for the details for each big idea. (Don't do anything else yet.) This should take about 1–2 minutes.
	2. Do not write anything yet! Look through the section for pictures, graphs, and tables. Take a minute to look at these—the author must have thought they were important. Also read over (but don't try to answer) the homework questions/problems at the end of the section. (For the visuals, the author must think these things illustrate something that is important enough to dedicate a significant amount of page real estate to it. For the homework problems, these illustrate what the author thinks you should be able to do once you know the content.) This process should take about 10–15 minutes.
	3. Actually read the text, one section at a time. For each section, jot down keywords and sentence fragments that remind you of the key ideas. You are not allowed to write more than the ¼ to ½ page allotted. (You don't need to write out the details—those are in the book, which you already have!) This process is time consuming, but shorter than what you're probably used to doing for your other teachers.
	4. Read the summary at the end of the chapter or section—this is what the author thinks you should know now that you've finished the reading. If there's anything you don't recognize, go back and look it up. This process should take about 5–10 minutes.
	You shouldn't need to use more than about one sheet of paper (both sides) per 10 pages of reading!

Taking Notes on Math Problems

Unit: Introduction

Details

MA Curriculum Frameworks (2016): SP5

Mastery Objective(s): (Students will be able to...)

• Take notes on math problems that both show and explain the steps.

Success Criteria:

- Notes show the order of the steps, from start to finish.
- A reason or explanation is indicated for each step.

Language Objectives:

• Be able to describe and explain the process of taking notes on a math problem.

Notes:

If you were to copy down a math problem and look at it a few days or weeks later, chances are you'll recognize the problem, but you won't remember how you solved it.

Solving a math problem is a process. For notes to be useful, they need to describe the process as it happens, not just the final result.

If you want to take good notes on how to solve a problem, you need your notes to show what you did at each step.

For example, consider the following problem:

How much heat is needed to increase the temperature of a 25 g sample of a metal with a specific heat capacity of $0.375 \frac{J}{g^{\circ}C}$ by 40 °C?

The process of solving this problem involves applying the equation $Q = mC\Delta T$, where Q is the amount of heat, m is the mass of the metal, C is the specific heat capacity of the metal, and ΔT is the temperature change. (Note that ΔT is only one quantity, even though it uses two symbols.)

Taking Notes on Math Problems

Details Unit: Introductio A good way to document the process is to use a T-chart, in which you show the steps of the solution on the left side, and you write an explanation of what you did and why for each step on the right side.

For this problem, your T-chart might look like the following:

Step	Description/Explanation
m = 25 g $C = 0.375 \frac{J}{g^{\circ}C}$ $\Delta T = 40 ^{\circ}C$ Q = quantity desired	Declare variables.
<u>Q</u> = <u>m C</u> ΔT	Choose a formula that gives the desired quantity. Make sure we have values for the other variables.
$Q = m C \Delta T$ Q = (25) (0.375) (40) Q = 375	Look up the values of any constants needed to solve the problem. Substitute for the variables and solve. (Show as much of the algebra as you think you'll need later.)
Q = 380 J	Round to the appropriate number of significant figures, include the units, and box the final answer.

You will notice that the answers are provided for many of the homework problems in these notes. This is because students are often unsure of whether they are doing a problem correctly until they see whether or not they got the correct answer. This means that if your teacher assigns these problems for homework, *it is not sufficient to just write down the answer*.

When a teacher says "show work," this does not necessarily mean you should show what <u>you</u> did to obtain the answer. Rather, it means:

- 1. Declare variables and assign them to values (with units).
- 2. Write down the relevant equation.
- 3. Substitute numbers for variables in the equation.
- 4. Solve for the missing variable.
- 5. Round to the appropriate number of significant figures and tack on the correct units.

This process would correspond to the left column of the above T-chart.

Use this space for summary and/or additional notes:

Big Ideas



Big Ideas

Chemistry

Unit: What Is Chemistry?

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Explain what chemistry is and what is studied in different branches of chemistry.

Success Criteria:

• Explanation describes what is studied in each of the branches of study described in this section.

Tier 2 Vocabulary: matter

Language Objectives:

• Understand and correctly use terms relating to each branch of chemistry.

Notes:

Details

<u>matter</u>: the "stuff" that everything is made of. Matter is anything that has mass and takes up space (has volume).

<u>chemistry</u>: the study of matter, its properties, how it behaves, how it's put together, and how it can be changed or rearranged .

<u>chemical</u>: a specific substance (regardless of size or shape) that has a specific arrangement of the atoms that it's made of, and has specific properties because of that arrangement.



Big Ideas	Details	Unit: What Is Chemistry?
	• chem	ical reactions
	0 d	ifferent ways atoms can rearrange (chemical reactions & equations)
	0 C	alculating how much of the reactants you use and products you make toichiometry)
	0 h	eat produced (or consumed) by chemical reactions
	0 h	ow fast chemicals react (kinetics)
	0 h	ow much chemicals react (equilibrium)
	0 a	cids & bases
		Branches of Chemistry
	The study o	f chemistry is divided into different branches, including:
	<u>organic che</u> contain	mistry: the study of chemicals and reactions involving molecules that carbon and hydrogen.
	<u>inorganic cl</u> do not	nemistry: the study of chemicals and reactions involving molecules that contain both carbon and hydrogen.
	biochemist process	<u>y</u> : the study of chemicals that play important roles in biological es, such as amino acids, lipids, and sugars.
	<u>physical che</u> include study o study o	emistry: the study of energy changes in chemistry. Some sub-fields thermodynamics (the study of heat energy), statistical mechanics (the f molecular collisions and momentum), and quantum mechanics (the f discrete energy changes at the sub-atomic level).
	<u>analytical c</u> chemic	nemistry: quantitative aspects of chemistry, such as determining what a a lis made of, how much of it reacts under certain conditions, <i>etc.</i>
	<u>green chem</u> used in	<u>iistry</u> : the study of making decisions about how chemicals are made or order to reduce the impact on the environment.

Unit: What Is Chemistry?

MA Curriculum Frameworks (2016): SP1, SP2, SP6, SP7

Mastery Objective(s): (Students will be able to...)

• Explain how the scientific method can be applied to a problem or question.

Success Criteria:

- Steps in a specific process are connected in consistent and logical ways.
- Explanation correctly uses appropriate vocabulary.

Tier 2 Vocabulary: theory, model, claim, law, peer

Language Objectives:

• Understand and correctly use terms relating to the scientific method, such as "peer review".

Notes:

The scientific method is a fancy name for "figure out what happens by trying it."

In the middle ages, "scientists" were called "philosophers." These were church scholars who decided what was "correct" by a combination of observing the world around them and then arguing and debating with each other about the mechanisms and causes.

During the Renaissance, scientists like Galileo Galilei and Leonardo da Vinci started using experiments instead of argument to decide what really happens in the world.







Big Ideas	Details Unit: What Is Chemistry?
	Before a claim is taken seriously, the original scientist and any others who are interested try everything they can think of to disprove the claim. If the claim holds up despite many attempts to disprove it, the claim gains support.
	peer review: the process by which scientists scrutinize, evaluate and attempt to disprove each other's claims.
	If a claim has gained widespread support among the scientific community and can be used to predict the outcomes of experiments (and it has <i>never</i> been disproven), it might eventually become a theory or a law.
	<u>theory</u> : a claim that has never been disproven, that gives an explanation for a set of observations, and that can be used to predict the outcomes of experiments.
	<u>model</u> : a way of viewing a set of concepts and their relationships to one another. A model is one type of theory.
	<u>law</u> : a claim that has never been disproven and that can be used to predict the outcomes of experiments, but that does not attempt to model or explain the observations.
	Note that the word "theory" in science has a different meaning from the word "theory" in everyday language. In science, a theory is a model that:
	 has never failed to explain a collection of related observations
	• <i>has never failed</i> to successfully predict the outcomes of related experiments
	For example, the theory of evolution <i>has never failed</i> to explain the process of changes in organisms caused by factors that affect the survivability of the species.
	If a repeatable experiment contradicts a theory, and the experiment passes the peer review process, the theory is deemed to be wrong. If the theory is wrong, it must either be modified to explain the new results, or discarded completely.

Big Ideas	Details	Unit: What Is Chemistry?
	Theories <i>vs.</i> Na	atural Laws
	The terms "theory" and "law" developed org definition of either term must acknowledge t outside of the scientific community, will not a definitions.	anically over many centuries, so any that common usage, both within and always be consistent with the
	Nevertheless, the following rules of thumb m	nay be useful:
	A <i>theory</i> is a model that attempts to explain simply describes or quantifies what happens explanation. Theories and laws can both be experiments.	<u>why</u> or <u>how</u> something happens. A <i>law</i> without attempting to provide an used to predict the outcomes of related
	For example, the <i>Law of Gravity</i> states th based on their masses and distances fror theory because the Law of Gravity does r other.	nat objects attract other objects m each other. It is a law and not a not explain <i>why</i> masses attract each
	Atomic Theory states that matter is made are themselves made up of smaller partic these particles are used to explain certain is a theory because we cannot see atoms the model gives an explanation for <i>why</i> s they do.	e of atoms, and that those atoms cles. The interactions between n properties of the substances. This s or prove that they exist. However, substances have the properties that
	A theory cannot become a law for the same a become a measurement, and a postulate car	reasons that a definition cannot nnot become a theorem.

	introduction. Laboratory	Page: A
Big Ideas	Details	Unit: Laborato
	Introduction: Laboratory	/
	Unit: Laboratory	
	Topics covered in this chapter:	
	Designing & Performing Experiments	26
	Laboratory Equipment	32
	Accuracy & Precision	35
	Uncertainty & Error Analysis	37
	Significant Figures	46
	Keeping a Laboratory Notebook	55
	Internal Laboratory Reports	59
	Formal Laboratory Reports	64
	Standards addressed in this chapter:	
	Massachusetts Curriculum Frameworks & Science Practices (2016):
	SP3: Planning & carrying out investigations	
	SP4: Analyzing & interpreting data	
	SP6: Constructing explanations	
	SP8: Obtaining, evaluating & communicating information	

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): SP1, SP3, SP8

Mastery Objective(s): (Students will be able to...)

• Create a plan and procedure to answer a question through experimentation.

Success Criteria:

- Experimental Design utilizes backward design.
- Experimental Design uses logical steps to connect the desired answer or quantity to quantities that can be observed or measured.
- Procedure gives enough detail to set up experiment.
- Procedure establishes values of control and independent variables.
- Procedure explains how to measure dependent variables.

Tier 2 Vocabulary: inquiry, independent, dependent, control

Language Objectives:

- Understand and correctly use the terms "dependent variable" and "independent variable."
- Understand and be able to describe the strategies presented in this section.

Notes:

Unlike many biology experiments in which the purpose is to observe something, most chemistry experiments involve observing changes to find out *what* happens, *under what conditions* it happens, and measuring *to what extent* it happens. Part of the purpose of chemistry experiments is to get practice making things happen, but another part of the purpose is for you to figure out *how* to get them to happen and how to measure them. This makes chemistry well-suited for teaching you how to design experiments.

The education "buzzword" for this is *inquiry-based experiments*, which means you (or your lab group) will need to figure out what to do to perform an experiment that answers a question about some aspect of chemistry. In this course, you will usually be given only an objective or goal and a general idea of how to go about achieving it. You and your lab group (with help) will decide the specifics of what to do, what to measure (and how to measure it), and how to make sure you are getting good results. This is a form of *guided* inquiry.

Big Ideas	Details Unit: Laboratory
	Framing Your Experiment
	Experiments are motivated by something you want to find out, observe, or calculate.
	Independent, Dependent, and Control Variables
	In an experiment, there is usually something you are doing, and something you are measuring or observing.
	independent variable: the conditions you are setting up. These are the predetermined values (the ones you pick). Because you choose the values, they are <i>independent</i> of what happens in the experiment. For example, if you are trying to figure out how much sugar you can dissolve in water at different temperatures, you are choosing the temperatures to test, so temperature is the <i>independent</i> variable.
	<u>dependent variable</u> : the things that happen in the experiment. These are the numbers you measure as a result of the experiment, which means they are <i>dependent</i> on what happens in the experiment. For example, if you are trying to figure out how much sugar you can dissolve in water, you don't know how much sugar dissolves until you measure it. This means the amount of sugar that dissolves is the <i>dependent</i> variable.
	<u>control variable</u> : things that you need to keep constant in the experiment. If there are a lot of factors that could affect your dependent variable, and you want to test the effect of one of them, you need to keep the other ones the same. Otherwise, you don't know what caused the changes that you are seeing.
	Control variables are usually parameters that could become independent variables in other experiments. For example, if you are trying to figure out how much sugar you can dissolve in water at different temperatures, you need to use the same volume (amount) of water each time. This means the volume of water that you use is a <i>control</i> variable.
	If someone asks what your independent, dependent and control variables are, the question simply means:
	 "What did you vary on purpose (independent variables)?"
	 "What did you measure (dependent variables)?"
	• "What did you keep the same for each trial (control variables)?"
	Use this space for summary and/or additional notes:

Big Ideas	Details Unit: Laboratory
	Qualitative Experiments
	If the goal of your experiment is to find out <i>whether or not</i> something happens at all, you need to set up a situation in which the phenomenon you want to observe can either happen or not, and then observe whether or not it does. The only hard part is making sure the conditions of your experiment don't bias whether the phenomenon happens or not.
	If you want to find out <i>under what conditions</i> something happens, what you're really testing is whether or not it happens under different sets of conditions that you can test. In this case, you need to test three situations:
	 A situation in which you are sure the thing will happen, to make sure you can observe it. This is your positive control.
	 A situation in which you sure the thing cannot happen, to make sure your experiment can produce a situation in which it doesn't happen and you can observe its absence. This is your negative control.
	3. A condition or situation that you want to test to see whether or not the thing happens. The condition is your independent variable, and whether or not the thing happens is your dependent variable.
	Quantitative Experiments
	If the goal of your experiment is to quantify (find a numerical relationship for) the extent to which something happens (the dependent variable), you need to figure out a set of conditions under which you can measure the thing that happens. Once you know that, you need to figure out how much you can change the parameter you want to test (the independent variable) and still be able to measure the result. This gives you the highest and lowest values of your independent variable. Then perform the experiment using a range of values for the independent value that cover the range from the lowest to the highest (or <i>vice-versa</i>).
	For quantitative experiments, a good rule of thumb is the 8 & 10 rule : you should have at least 8 data points, and the range from the highest to the lowest values tested should span at least a factor of 10.

Details **Big Ideas** Unit: Laboratory Letting the Chemistry Design the Experiment Determining what to measure usually means determining what you need to know and working backwards to figure out how to get there from quantities that you can measure. Especially in chemistry, most of the complexity of experiments comes from the fact that there are few quantities that we can measure directly. When we have to measure quantities indirectly, we need to be clever, and we need to think about all of the possible ways we could end up with an invalid result, so we can design the appropriate safeguards into the experiment. For a quantitative experiment, it's often best to start with a mathematical formula that includes the quantity you want to determine. Then, you need to find the values of the other quantities in the equation, either by measuring them directly or by performing an experiment in which you can determine them by measuring other things. As an illustrative example, suppose you want to calculate the heat released per mole of reactant in a specific chemical reaction. (Don't worry if you don't understand what this means—we will get there later in the course.) You might go through the following thought process: 1. We can't measure heat directly, so we need an equation to calculate it. We can use the equation: $Q = mC\Delta T$ In order to use this equation to calculate heat (Q), we need to find m (mass), C (specific heat capacity), and ΔT (temperature change), which means we need to use the heat to change the temperature of a mass of something. An experimental device that does this is a bomb calorimeter^{*}, so we will use one of those. ^{*} A "bomb calorimeter" is a device that uses the heat of a chemical reaction to heat up a known amount of water so you can measure how much heat is produced. Despite the cool-sounding name, a bomb calorimeter doesn't actually involve blowing anything up.

Designing & Performing Experiments

Big Ideas	Details	0 0	0 1	Unit: Laboratory		
	2. Dete	ermine our inde	pendent variables in the	above equation.		
	 We predetermine <i>m</i> (the mass of the water in the calorimeter) by putting a known mass of water into the calorimeter. We can measure the mass of the water using a balance. 					
	b.	C (the specific constant that v	heat capacity of the wate we can look up.	er in the calorimeter) is a		
	3. Determine our dependent variables in the equation.					
	•	We need to ca means we nee the end, and s thermometer.	lculate ΔT (the temperated to measure the temperated to measure the temperated to measure the temperated by the can measure of the temperated by the te	ure change of the water). This rature at the beginning and at e the temperatures with a		
	4. Determine our control variables.					
	a.	We want the heat, we need much of the ch mass of the ch	heat per mole of reactant to control the number of nemical we use), which m emical and convert it to r	 If we're going to calculate the f moles of reactant (<i>i.e.</i>, how leans we need to measure the moles. 		
	b.	We need to m amounts of ev	ake sure all of the chemic erything else.	cal reacts, so we need use excess		
	C.	We need to m we need to set chemicals are	ake sure no heat is lost to t the experiment up so w already inside the calorin	o the environment. This means e can start the reaction after the neter.		
	Notice that wanted to problems v	our entire expe know at the end ve need to avoid	eriment is ultimately dete d of it, and figuring out w d.	ermined by starting with what we hat makes it happen, and what		

Details

The Experimental Design Process

- 1. **Figure out what you want to know.** Decide on an experiment that you could use to find it. If what you want to know is a term in an equation, your experiment will involve finding out values for the other variables in the equation.
- 2. Determine your independent variables. Based on your experiment and the equation that goes with it if there is one, determine your independent variables. Figure out what they are, how you are going to determine/measure them, and which values you are going to choose for them. (You may want to wait to choose the values until you know what your dependent and control variables are, and use the expected values to backward calculate your choices for your independent variables.)
- 3. **Determine your dependent variables.** Figure out what they are, and how *and <u>when</u>* you are going to measure them.
- 4. **Determine your control variables.** Think about the things that you already know that you need to keep constant. Then start thinking about what could go wrong, which will lead you to other things you need to keep constant.
- 5. Set up your experiment and do a test run. Use your test run to make sure you can actually measure what you think you can measure and make sure you are getting results that make sense. *This means you need to perform the calculations for your test run before doing the rest of the experiment,* in case you need to modify your procedure. You will be sad if you finish your experiment and go home, only to find out at 2:00 am the night before the write-up is due that it didn't work.

More complex experiments use this same process, except that each step might have several sub-steps, and/or that each step or sub-step might be its own completely separate experiment.

Laboratory Equipment

Unit: Laboratory

MA Curriculum Frameworks (2016): SP3

Mastery Objective(s): (Students will be able to...)

• Recognize and identify common laboratory equipment used in chemistry.

Success Criteria:

• Be able to give the name and describe the use for each of the pieces of equipment in this section.

Tier 2 Vocabulary: graduated, ring, stand

Language Objectives:

• Identify laboratory equipment by name and explain what it is used for.

Notes:

The following are some of the common pieces of laboratory apparatus.

Description	Picture	Description	Picture
beaker Used as a cup/container. May be heated. Volume markings are approximate.		graduated cylinder Used for accurate measurement of liquid volume.	(Junnanana)
Erlenmeyer flask Used as a container. Contents may be swirled to mix. May be heated.	\square	test tube Used for mixing and reacting small quantities of liquids.	
Bunsen burner Used to heat chemicals in beakers, flasks or test tubes.		ring stand Used to support lab apparatus.	

Laboratory Equipment

Page: 33 Unit: Laboratory

Details Unit: Laborato					
Description	Picture	Description	Picture		
tongs Used for picking up and holding hot things.	Y	test tube holder Used for holding a test tube while heating.	60		
clamp		pinch clamp			
Used to attach a piece of equipment to a ring stand.		Used to clamp rubber tubing so no liquid comes out.	Z		
wire gauze		iron ring			
Used to spread the heat of a flame onto the bottom of a beaker or flask.		Fastens to ring stand. Used to hold lab apparatus.	C AND		
mortar & pestle		evaporating dish			
Used to grind chemicals to a fine powder.	O	Used to hold solutions over a Bunsen burner to evaporate the liquid.	\bigcirc		
stirring rod		pipette	Ω		
Used to stir mixtures and to help pour liquids without spilling.		Used to transfer small amounts of liquid.			
watch glass		crucible & cover			
Used to cover liquids being heated in a beaker to avoid spatter. Should not be heated.	\bigcirc	Used to heat small amounts of solids to high temperatures.	ð		
funnel	\bigcirc	pipestem triangle			
Used to help pour liquids without spilling and/or to hold filter paper.	Y	Used to hold a crucible while heating.			

Use this space for summary and/or additional notes:

Big Ideas

Laboratory Equipment



Accuracy & Precision

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): SP3

Mastery Objective(s): (Students will be able to...)

• Correctly use the terms "accuracy" and "precision" in a scientific context.

Success Criteria:

- Be able to give definitions for "accuracy" and "precision."
- Be able to recognize situations as accurate/inaccurate and/or precise/imprecise.

Tier 2 Vocabulary: accurate, precise

Language Objectives:

• Be able to describe the difference between accuracy and precision.

Notes:

Science relies on making and interpreting measurements, and the accuracy and precision of these measurements affect what you can conclude from them.

Random vs. Systematic Errors

Random errors are natural uncertainties in measurements because of the limits of precision of the equipment used. Random errors are assumed to be distributed around the actual value, without bias in either direction. Systematic errors occur from specific problems in your equipment or your procedure. Systematic errors are often biased in one direction more than another, and can be difficult to identify.

Accuracy vs. Precision

The words "accuracy" and "precision" have specific meanings in science.

- <u>accuracy</u>: for a single measurement, how close the measurement is to the "correct" or accepted value. For a group of measurements, how close the <u>average</u> is to the accepted value.
- <u>precision</u>: for a single measurement, how finely the measurement was made. (How many decimal places it was measured to.) For a group of measurements, how close the measurements are to each other.

Accuracy & Precision


Uncertainty & Error Analysis

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): SP4

Mastery Objective(s): (Students will be able to...)

• Determine the uncertainty of a measured or calculated value.

Success Criteria:

- Take analog measurements to one extra digit of precision.
- Correctly estimate measurement uncertainty.
- Correctly read and interpret stated uncertainty values.
- Correctly propagate uncertainty through calculations involving addition/subtraction and multiplication/division.

Tier 2 Vocabulary: uncertainty, error

Language Objectives:

- Understand and correctly use the terms "uncertainty" and "relative error."
- Correctly explain the process of estimating and propagating uncertainty.

Notes:

In science, unlike mathematics, there is no such thing as an exact answer. Ultimately, every quantity is limited by the precision and accuracy of the measurements that it came from. If you can only measure a quantity to within 10 %, that means any calculation that is derived from that measurement can't be any better than ± 10 %.

Error analysis is the practice of determining and communicating the causes and extents of uncertainty in your results. Error analysis involves understanding and following the uncertainty in your data, from the initial measurements to the final calculated and reported results.

Note that the word "error" in science has a different meaning from the word "error" in everyday language. In science, "error" means "uncertainty." If you report that you drive (2.4 ± 0.1) miles to school every day, you would say that this distance has an error of ± 0.1 mile. This does not mean your car's odometer is wrong; it means that the actual distance *could be* 0.1 mile more or 0.1 mile less—*i.e.*, somewhere between 2.3 and 2.5 miles. When you are analyzing your results, <u>never</u> use the word "error" to mean mistakes that you might have made!

Uncertainty & Error Analysis

Big Ideas	Details Unit: Laboratory
	Uncertainty
	The uncertainty or error of a measurement describes how close the actual value is likely to be to the measured value. For example, if a length was measured to be 22.34 cm, and the uncertainty was 0.31 cm (meaning that the measurement is only known to within ± 0.31 cm), we could represent this measurement in either of two ways:
	(22.34 ± 0.31) cm 22.34(31) cm
	The first of these states the variation (±) explicitly in cm (the actual unit). The second is shows the variation in the last digits shown.
	What it means is that the true length is approximately 22.34 cm, and is statistically likely [*] to be somewhere between 22.03 cm and 22.65 cm.
	Absolute Error
	Absolute error (or absolute uncertainty) refers to the uncertainty in the actual measurement. For the measurement (22.34 \pm 0.31) cm, the absolute error is \pm 0.31 cm.
	Relative Error
	Relative error shows the error or uncertainty as a fraction of the total.
	The formula for relative error is R.E. = $\frac{\text{uncertainty}}{\text{measured value}}$
	For the measurement (22.34 \pm 0.31) cm, the relative error would be 0.31 out of 22.34. Mathematically, we express this as:
	$R.E. = \frac{0.31}{22.34} = 0.0139$
	Note that relative error is dimensionless (does not have any units). This is because the numerator and denominator have the same units, so the units cancel.
	Percent Error
	Percent error is relative error expressed as a percentage. You can turn relative error into percent error by multiplying by 100.
	In the example above, the relative error of 0.0139 would be 1.39 % error.
	[*] Statistically, the uncertainty is one standard deviation. <i>I.e.,</i> if multiple measurements are taken, approximately two-thirds of those measurements will lie within the uncertainty (plus or minus) of the stated value.

Uncertainty of Measurements

If you have the ability to measure a quantity that is not changing (such as the mass or length of an object), you will get the same value every time you measure it. This means you have only one data point.

When you have only one data point, the uncertainty is the limit of how well you can measure it. This will be your best educated guess, based on how closely you think you actually measured the quantity. This means you need to take measurements as carefully and precisely as possible, because *every careless measurement needlessly increases the uncertainty of the result*.

Digital Measurements

Details

For digital equipment, if the reading is <u>stable</u> (not changing), look up the published precision of the instrument in its user's manual. (For example, many balances used in high schools have a readability of 0.01 g but are only precise to within \pm 0.02 g.) If there is no published value (or the manual is not available), assume the uncertainty is \pm 1 in the last digit.

Analog Measurements

When making analog measurements, <u>always</u> estimate one extra digit beyond the finest markings on the equipment. For example, in the diagram below, the graduated cylinder is marked in 1 mL increments. When measuring volume in this graduated cylinder, you would estimate and write down the volume to the nearest 0.1 mL, as shown:



In the above experiment, you should record the volume as 32.0 ± 0.1 mL. It would be <u>inadequate</u> to write the volume as 32 mL; you *must* write 32.0 mL, or better yet, (32.0 ± 0.1) mL

The zero at the end of 32.0 mL is <u>not</u> extra. It is necessary to show that *you* measured the volume to the nearest tenth, not to the nearest one.

Use this space for summary and/or additional notes:

Big Ideas

Uncertainty & Error Analysis

Big Ideas

Details

When estimating, the uncertainty depends on how well you can see the markings, but you can usually assume that the estimated digit has an uncertainty of $\pm \frac{1}{10}$ of the finest markings on the equipment. Here are some examples:

Equipment	Markings	Estimate To	Assumed Uncertainty
ruler	1 mm	0.1 mm	± 0.1 mm
25 mL graduated cylinder	0.2 mL	0.02 mL	± 0.02 mL
thermometer	1 °C	0.1 °C	± 0.1 °C

Homework Problems

Write the readings that you would record (estimated to one extra decimal place) and the assumed uncertainty for each of the following thermometers and graduated cylinders.



Propagating Uncertainty in Calculations

When you perform calculations using numbers that have uncertainty, you need to propagate the uncertainty through the calculation.

Addition & Subtraction

When quantities with uncertainties are added or subtracted, add the quantities to get the answer, then add the uncertainties to get the total uncertainty.

Sample Problem:

Details

Big Ideas

- Q: A substance is being heated. You record the initial temperature as (23 ± 0.2) °C, and the final temperature as (84 ± 0.2) °C. You need to calculate the temperature change (ΔT) with its uncertainty to use in a later calculation. What is the temperature change?
- A: To calculate ΔT , simply subtract:

$$\Delta T = T_{final} - T_{initial} = 84 - 23 = 61^{\circ}\text{C}$$

To calculate the uncertainty, add the individual uncertainties (even though the quantities were subtracted):

u = 0.2 + 0.2 = 0.4 °C

Report the value as: $\Delta T = (61 \pm 0.4)^{\circ}C$

Multiplication & Division

Because most calculations that we will perform in chemistry involve multiplication and/or division, you can

For calculations involving multiplication and division, estimate the uncertainty of your calculated answer by adding the relative errors and applying the total relative error to your result.

- 1. Perform the calculation for the desired quantity.
- 2. Divide the uncertainty (the ±) for each quantity by its measured value to determine its relative error.

- 3. Add up all of the relative errors to get the total relative error.
- 4. Multiply your calculated result by the total relative error to get its uncertainty (the ± amount).

Uncertainty & Error Analysis

	-
tails	Unit: Laboratory
te: Most of the calculations that you will perform in chemistry Itiplication and/or division, so almost all of your uncertainty co oughout the course will use relative error.	v involve alculations
mple Problem #1:	
You want to determine the amount of heat released by a cheme You use the heat from the reaction to heat up some water in a container called a "bomb calorimeter". You will calculate the h equation: $Q = mC\Delta T$.	nical reaction. In insulated heat using the
Suppose you recorded the following data (including uncertaint	ies):
• The mass of the water in the calorimeter is (24.8 \pm 0.1) g.	
 The temperature change of the water was (12.4± 0.2) °C. 	
• The specific heat capacity of water is $4.181 \frac{1}{g \cdot C}$. (This is a p	ublished value.
The uncertainty of this value is so small that we can leave it calculations.)	out of our
The heat released by the reaction is given by the equation:	
$Q = mC\Delta T$	
Q = (24.8)(4.181)(12.4)	
<i>Q</i> = 1 285.74 J	
The relative errors for the two quantities that we measured an	re:
• mass: $\frac{0.1}{24.8} = 0.00403$	
• temperature change: $\frac{0.2}{12.4} = 0.01613$	
The total relative error is 0.00403 + 0.01613 = 0.02016	
The uncertainty is therefore $(0.02016)(1285.74) = \pm 25.92 \text{ J}$	
(Note that the absolute uncertainty has the same units as the	measurement.)
We would report the measurement as (1 285.74 \pm 25.92) J.	
	the image of the calculations that you will perform in chemistry triplication and/or division, so almost all of your uncertainty of oughout the course will use relative error. mple Problem #1: You want to determine the amount of heat released by a chemination of the transmitter of the trans

Big Ideas	Details Unit: Laborato
	Rounding
	In the example above, the uncertainty tells us that our actual result could be
	different from our calculated value by as much as 25.92 J.
	However, we only estimated one digit (which happened to be the tenths place) when we took our measurements. This means we have only one digit of uncertainty. Because we can't report more precision than we actually have, we need to round the calculated uncertainty off, so that we have only one unrounded digit. This means we should report our uncertainty as ± 30 J.
	It wouldn't make sense to report our answer as (1285.74 ± 30) J. Think about that if the <i>tens</i> digit could be different from our calculated value, there is no point in reporting the ones or tenths digits. So we need to round our calculated answer to the same place value as the uncertainty—the tens place.
	This means our final, rounded answer should be (1 290 ± 30) J.

Big Ideas	Details Unit: Laboratory
	Sample Problem #2:
	 Q: You need to find the density of a piece of metal. We measure its mass on a balance to be (24.75 ± 0.02) g. You measure its volume in a graduated cylinder using water displacement, and you find the volume to be (7.2 ± 0.1) mL. Calculate the density, including its uncertainty.
	A: 1. Calculate the density.
	$\rho = \frac{m}{V} = \frac{24.75 \text{ g}}{7.2 \text{ mL}} = 3.4375 \frac{\text{g}}{\text{mL}}$
	2. Calculate the relative errors of your two measurements:
	$R.E{mass} = \frac{\text{uncertainty}}{\text{measured value}} = \frac{0.02}{24.75} = 0.000808$
	$R.E_{\text{volume}} = \frac{0.1}{7.2} = 0.013889$
	3. Add the individual relative errors together to get the total R.E.:
	0.000808 + 0.013889 = 0.014697
	4. Multiply the total R.E. by the density to get the uncertainty:
	$3.4375 \times 0.014697 = 0.050521$
	Because you only estimated one decimal place of uncertainty, you need to round the uncertainty off to \pm 0.05.
	Because uncertainty is rounded to the hundredths place, you need to also round your answer to the hundredths place:
	$ ho = (3.44 \pm 0.05) \frac{g}{mL}$

Uncertainty & Error Analysis

Big Ideas	Details Unit: Laborato		
	Homework Problems		
	Because the answers are provided, you must show sufficient work in order to receive credit.		
	1. You need to combine three liquids and calculate the total volume. You measured the individual volumes as (12.36 ± 0.02) mL, (37.4 ± 0.2) mL, and (61.0 ± 0.1) mL. What is the total volume, including the uncertainty?		
	Answer: (110.8 ± 0.3) mL		
	 A sample of (0.517 ± 0.008) moles of a chemical is dissolved to make (1.362 ± 0.005) liters of solution. 		
	 What is the concentration in moles per liter? (Divide the amount in moles by the volume in liters.) 		
	Answer: 0.3796 ^{mol} (Don't worry about rounding yet.) b. What are the relative errors of the number of moles and the number of liters? What is the total relative error?		
	 Answers: moles: R.E. = 0.015 volume: R.E. = 0.0037 total R.E. = 0.019 c. Calculate the uncertainty of the concentration of the solution and express your answer as the concentration (from part a above) plus or minus the uncertainty that you just calculated, with correct rounding. 		
	Answer: $(0.380 \pm 0.007) \frac{mol}{L}$		

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

- Identify the significant figures in a number.
- Perform calculations and round the answer to the appropriate number of significant figures

Success Criteria:

- Be able to identify which digits in a number are significant.
- Be able to count the number of significant figures in a number.
- Be able to determine which places values will be significant in the answer when adding or subtracting.
- Be able to determine which digits will be significant in the answer when multiplying or dividing.
- Be able to round a calculated answer to the appropriate number of significant figures.
- Tier 2 Vocabulary: significant, round

Language Objectives:

• Explain the concepts of significant figures and rounding.

Notes:

Because it would be tedious to calculate the uncertainty for error for every calculation in chemistry, we often use significant figures (or significant digits) as a simple way to estimate and represent the uncertainty.

Significant figures are based on the following approximations:

- All stated values are rounded off so that the uncertainty is only in the last unrounded digit.
- Assume that the uncertainty in the last unrounded digit is ±1.
- The results of calculations are rounded so that the uncertainty of the result is only in the last unrounded digit and is assumed to be ±1.

Big Ideas	Details Unit: Laboratory
	Note that using significant figures gives less information than stating the measurement with its uncertainty. This is why, when you take measurements and perform calculations in the laboratory, you will estimate the actual uncertainty of each measurement and calculate the uncertainty of your results. However, for homework problems and written tests, you will use significant figures as a simple way to keep track of the approximate effects of uncertainty on your answers.
	In the example on page 42, we rounded the number 1285.74 off to the tens place resulting in the value of 1290, because we couldn't show more precision than we actually had.
	In the number 1290, we would say that the first three digits are "significant", meaning that they are the part of the number that is not rounded off. The zero in the ones place is "insignificant," because the digit that was there was lost when we rounded.
	<u>significant figures</u> (significant digits): the digits in a measured value or calculated result that are not rounded off. (Note that the terms "significant figures" and "significant digits" are used interchangeably.)
	insignificant figures: the digits in a measured value or calculated result that were "lost" (became zeroes before a decimal point or were cut off after a decimal point) due to rounding.

Big Ideas	Details	Unit: Laboratory		
	Identifying the Significant Digits in a NumberThe first significant digit is where the "measured" part of the number begins—the first digit that is not zero.The last significant digit is the last "measured" digit—the last digit whose true value is known.If the number doesn't have a decimal point, the last significant digit will be the last digit that is not zero. (Anything after that has been rounded off.)			
	Example: If we round the number 2345 get 235000. (Note that because the dig was 5 or greater, so we had to "round u first three digits (the 2, 3, and 5) are the digits (the zeroes at the end) are the ins	hple: If we round the number 234567 to the thousands place, we would 235000. (Note that because the digit after the "4" in the thousands place 5 or greater, so we had to "round up".) In the rounded-off number, the three digits (the 2, 3, and 5) are the significant digits, and the last three is (the zeroes at the end) are the insignificant digit will be the last digit vn. (Anything rounded after the decimal point gets chopped off.) hple: If we round the number 11.223344 to the hundredths place, it id become 11.22. When we rounded the number off, we "chopped off" extra digits. e number is in scientific notation, it has a decimal point. Therefore, the re rules tell us (correctly) that all of the digits before the "times" sign are ficant.		
	 If the number has a decimal point, the last shown. (Anything rounded after the decimal point) 			
	Example: If we round the number 11.22 would become 11.22. When we rounder the extra digits.			
	 If the number is in scientific notation, it above rules tell us (correctly) that all of significant. 			
	In the following numbers, the significant figur	es have been underlined:		
	• <u>13</u> 000	• <u>6804.30500</u>		
	• 0.0 <u>275</u>	• 6.0×10^{23}		
	• 0.0 <u>150</u>	 <u>3400.</u> (note the decimal point at the end) 		
	L	· · · · · · · · · · · · · · · · · · ·		

Mathematical Operations with Significant Figures

Addition & Subtraction

When adding or subtracting, calculate the total normally. Then identify the smallest place value where nothing is rounded. Round your answer to that place.

For example, consider the following problem.

Details

Big Ideas

<u>pro</u>	blem:		<u>"sig figs" equivalent:</u>
+	123000 0.0075 <u>1650</u> 124650.0075 ▲	± 1000 ± 0.0001 _± 10 ± 1010.0001	123 ???.??? ? 0.007 5 <u>+ 165?.??? ?</u> 124 ???.??? ? (Check this digit for rounding)

In the first number (123 000), the hundreds, tens, and ones digit are zeros, presumably because the number was rounded to the nearest 1000. The second number (0.0075) is presumably rounded to the ten-thousandths place, and the number 1650 is presumably rounded to the tens place.

The first number has the largest uncertainty, so we need to round our answer to the thousands place to match, giving $125\ 000\ \pm\ 1\ 000$.

A silly (but correct) example of addition with significant digits is:

100 + 37 = 100

Multiplication and Division

When multiplying or dividing, calculate the result normally. Then count the total number of significant digits in the values that you used in the calculation. Round your answer so that it has the same number of significant digits as the value that had the *fewest*.

Consider the problem:

The answer (without taking significant digits into account) is $34.52 \times 1.4 = 48.328$

The number 1.4 has only two significant digits, so we need to round our answer so that it also has only two significant digits. This means we should round our answer to 48.

A silly (but correct) example of addition with significant digits is:

 $234 \times 1 = 200$

Big Ideas	Details	Unit: Laboratory
	Mixed Operations	
	For mixed operations, keep all of the digits until you're finished don't accumulate), but keep track of the last significant digit in a line over it (even if it's not a zero). Once you have your final the correct number of significant digits. Don't forget to use the operations (PEMDAS)!	I (so round-off errors each step by putting answer, round it to correct order of
	For example:	
	137.4×52+120×1.77	
	$(137.4 \times 52) + (120 \times 1.77)$	
	7,144.8+212.4=7,357.2=7,400	
	Note that in the above example, we kept all of the digits until avoid introducing small rounding errors at each step, which car change the final answer. Notice how, if we had rounded off the step, we would have gotten the wrong answer:	the end. This is to n add up to enough to e numbers at each
	$137.4 \times 52 + 120 \times 1.77$ (137.4 \times 52) + (120 \times 1.77) 7, \overline{10} = 7, \overline{3}10 = 7, 300 \checkmark	

What to Do When Rounding Doesn't Give the Correct Number of Significant Figures

If you have a different number of significant digits from what the rounding shows, you can place a line over the last significant digit, or you can place the whole number in scientific notation. Both of the following have <u>four</u> significant digits, and both are equivalent to writing $13,000 \pm 10$

13000

Details

• 1.300×10⁴

When Not to Use Significant Figures

Significant figure rules only apply in situations where the numbers you are working with have a limited precision. This is usually the case when the numbers represent measurements. <u>Exact</u> numbers have infinite precision, and therefore have an infinite number of significant figures. Some examples of exact numbers are:

- Pure numbers, such as the ones you encounter in math class.
- Anything you can count. (*E.g.*, there are 24 people in the room. That means exactly 24 people, not 24.0 ± 0.1 people.)
- Whole-number exponents in formulas. (*E.g.*, the area of a circle is πr^2 . The exponent "2" is a pure number.)

You should also avoid significant figures any time the uncertainty is likely to be substantially different from what would be implied by the rules for significant figures, or any time you need to quantify the uncertainty more exactly.

Summary

Significant figures are a source of ongoing stress among chemistry students. To make matters simple, realize that few formulas in chemistry involve addition or subtraction, so you can usually just apply the rules for multiplication and division: look at each of the numbers you were given in the problem. Find the one that has the fewest significant figures, and round your final answer to the same number of significant figures.

If you have absolutely no clue what else to do, round to three significant figures. You would have to measure quite carefully to have more than three significant figures in your original data, and three is usually enough significant figures to avoid unintended loss of precision, at least in a high school chemistry course. ©

Big Ideas	Details		Unit: Laboratory
	Homework Problems		
	 For each of the fo Write the assume 	ollowing, Underline the significant figures i ed uncertainty as ± the appropriate quanti	n the number and ty.
	<u>573</u> 00 ± 100 ← 9	Sample problem with correct answer.	
	a. 13500		
	b. 26.0012		
	c. 01902		
	d. 0.000000	025	
	e. 320.		
	f. 6.0 × 10 ⁻⁷		
	g. 150.00		
	h. 10		
	i. 0.005 310 0	0	

Big Ideas	Details	U U	U	Unit: Laboratory
	2. Round signifi	d off each of the fo icant digit if neces	ollowing numbers as inc sary.	licated and indicate the last
	a.	13500 to the ne	arest 1000	
	b.	26.0012 to the n	earest 0.1	
	C.	1902 to the near	est 10	
	d.	0.000025 to the	nearest 0.000 01	
	e.	320. to the near	est 10	
	f.	6.0×10^{-7} to the	nearest 10 ⁻⁶	
	g.	150.00 to the ne	arest 100	
	h.	10 to the neares	t 100	

Big Ideas	Details	Ũ	U		Unit: Laboratory
	3. Solve num	e the following mat ber of significant fig	h problems and rou gures.	nd your answer t	o the appropriate
	a.	3 521 × 220			
	b.	13 580.160 ÷ 113	3		
	c.	2.71828 + 22.4 -	8.31 - 62.4		
	d.	23.5 + 0.87 × 6.0	2 – 105 (Remember	r PEMDAS!)	

Keeping a Laboratory Notebook

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): SP3, SP8

Mastery Objective(s): (Students will be able to...)

- Determine which information to record in a laboratory notebook.
- Record information in a laboratory notebook according to practices used in industry.

Success Criteria:

- Record data accurately and correctly, with units and including estimated digits.
- Use the correct protocol for correcting mistakes.

Language Objectives:

• Understand and be able to describe the process for recording lab procedures and data.

Notes:

A laboratory notebook serves two important purposes:

- 1. It is a diary of what you did in case you want to look up the details later.
- 2. It is a legal record of what you did and when you did it.

In a research laboratory, you would normally do a write-up in your lab notebook whenever you do a significant experiment that you believe you might want to refer back to sometime in the future.

Your Notebook as an Official Record

Laboratory notebooks are kept by scientists in research laboratories and high tech companies. If a company or research institution needs to prove that you did a particular experiment on a particular date and got a particular set of results (perhaps to apply for a patent), your lab notebook is the primary evidence. This means you need to maintain your lab notebook in a way that gives it the best chance of being able to prove beyond a reasonable doubt exactly what you did and exactly when you did it.

Keeping a Laboratory Notebook

Big Ideas	Details	Unit: Laboratory
	For companies that use laboratory notebooks in this way, there are guidelines that exist to prevent mistakes that could compromise the notebook. Details may vary somewhat from one company to anoth probably similar to these, and the spirit of the rules is the same.	a set of e integrity of the er, but are
	 All entries in a lab notebook must be hand-written in ink. 	
	 Your actual procedure and all data must be recorded directly notebook, not recorded elsewhere and copied in. 	into the
	 All pages must be numbered consecutively, to show that no p removed. If your notebook did not come with pre-numbered to write the page number on each page before using it. 	ages have been pages, you need
	 Start each experiment on a new page. 	
	 Sign and date the bottom of the each page when you finish reinformation on it. (In industry, each page needs to be witnesselse, usually your supervisor. The date that an entry is consid happened is the date it was witnessed, even if that is much la when it was originally recorded.) 	ecording sed by someone ered to have ter than the date
	 When crossing out an incorrect entry in a lab notebook, never Always cross it out with a single line through it, so that it is sti read the original mistake. (This is to prove that it was a mista didn't change your data or observations.) Any time you cross write your initials and the date next to the change. 	r obliterate it. Il possible to ke, and you something out,
	 Never remove pages from a lab notebook for any reason. If yo out an entire page, you may do so with a single large "X". If yo a brief explanation of why you crossed out the page, and sign cross-out. 	ou need to cross ou do this, write and date the
	 Never, ever change data after the experiment is completed. I data, right or wrong, is what you actually observed. Changing constitutes fraud, which is a form of cheating that is every bit plagiarism. 	Really. Your ; your data as bad as
	 Never change anything on a page you have already signed and realize that an experiment was flawed, leave the bad data wh a note that says "See page" with your initials and date r addendum. On the new page, refer back to the page number and describe briefly what was wrong with it. Then, give the c information and sign and date it as you would an experiment. 	d dated. If you ere it is and add next to the of the bad data orrect
	 Never, ever erase or cover with white-out anything in a lab no or covered-up data is considered the same as faked or change scientific community. 	otebook. Erased ed data in the

Use this space for summary and/or additional notes:

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Keeping a Laboratory Notebook Page: 57 Unit: Laboratory

Big Ideas	Details Unit: Laboratory
	Recording Data
	Here are some general rules for working with data. (Most of these are courtesy of Dr. John Denker, at http://www.av8n.com/physics/uncertainty.htm):
	• Write something about what you did on the same page as the data, even if it is a very rough outline. Your procedure notes should not get in the way of actually performing the experiment, but there should be enough information to corroborate the detailed summary of the procedure that you will write afterwards. (Also, for evidence's sake, the sooner after the experiment that you write the detailed summary, the more weight it will carry in court.)
	• Keep <u>all</u> of the raw data, whether you will use it or not.
	• Don't discard a measurement, even if you think it is wrong. Record it anyway and put a "?" next to it. You can always choose not to use the data point in your calculations (as long as you give an explanation).
	 Never erase or delete a measurement. The only time you should ever cross out recorded data is if you accidentally wrote down the wrong number.
	 Record all digits. Never round off original data measurements. If the last digit is a zero, you must record it anyway!
	• For analog readings (<i>e.g.,</i> ruler, graduated cylinder, thermometer), always estimate and record one extra digit.
	Always write down the units with each measurement!
	• Record <i>every</i> quantity that will be used in a calculation, whether it is changing or not.
	• Don't convert in your head before writing down a measurement. Record the original data in the units you actually measured it in, and convert in a separate step.
	Calculations
	• Use enough digits to avoid unintended loss of significance. (Don't introduce round-off errors in the middle of a calculation.) This usually means use at least two more digits than the number of "significant figures" you expect your answer to have.
	 Use few enough digits to be reasonably convenient.
	 Record uncertainty separately from the measurement. (Don't rely on "sig figs" to express uncertainty.)
	• Leave digits in the calculator between steps. (Don't round until the end.)
	• When in doubt, keep plenty of "guard digits" (digits after the place where you think you will end up rounding).
	Use this space for summary and/or additional notes:

Big Ideas	Details Unit: Laboratory
	Integrity of Data
	Your data are your data. In classroom settings, people often get the idea that the goal is to report an uncertainty that reflects the difference between the measured value and the "correct" value. That idea certainly doesn't work in real life—if you knew the "correct" value you wouldn't need to make measurements!
	In all cases—in the classroom and in real life—you need to determine the uncertainty of your own measurement by scrutinizing your own measurement procedures and your own analysis. Then you judge how well they agree.
	For example, we would say that the quantities 10 ± 2 and 11 ± 2 agree reasonably well, because there is considerable overlap between their probability distributions. However, 10 ± 0.2 does not agree with 11 ± 0.2 , because there is no overlap.
	If you get an impossible result or if your results disagree with well-established results, you should look for and comment on possible problems with your procedure and/or measurements that could have caused the differences you observed. You must <i>never</i> fudge your data to improve the agreement.
	Your Laboratory Notebook is <i>Not</i> a Report
	Many high school students are taught that a laboratory notebook should be a journal-style book in which they must write perfect after-the-fact reports, but they are not allowed to change anything if they make a mistake. This is not at all what laboratory notebooks were ever meant to be. A laboratory notebook does not need to be anything more than an official signed and dated record of your procedure (what you did) and your data (what happened) at the exact instant that you took it and wrote it down.
	Of course, because it is your journal, your laboratory notebook <i>may</i> contain anything else that you think is relevant. You may choose to include an explanation of the motivations for one or more experiments, the reasons you chose the procedure that you used, alternative procedures or experiments you may have considered, ideas for future experiments, <i>etc.</i> Or you may choose to record these things separately and cross-reference them to specific pages in your lab notebook.
	Use this space for summary and/or additional notes:

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): SP3, SP8

Mastery Objective(s): (Students will be able to...)

• Write an internal laboratory report that appropriately communicates all of the necessary information.

Success Criteria:

- The report has the correct sections in the correct order.
- Each section contains the appropriate information.

Language Objectives:

- Understand and be able to describe the sections of an internal laboratory report, and which information goes in each section.
- Write an internal laboratory report with the correct information in each section.

Notes:

An internal laboratory report is written for co-workers, your boss, and other people in the company or research facility that you work for. It is usually a company confidential document that is shared internally, but not shared outside the company or facility. Every lab you work in, whether in high school, college, research, or industry, will have its own preferred internal report format.

It is much more important to understand what *kinds* of information you need to report and what you will use it for than it is to get attached to any one format. The format we will use in this class is based on the outline of the actual experiment.

Title & Date

Each experiment should have the title and date the experiment was performed written at the top. The title should be a descriptive sentence fragment (usually without a verb) that gives some information about the purpose of the experiment.

Objective

This should be a one or two-sentence description of what you are trying to determine or calculate by performing the experiment.

Details Unit: Laboratory
Experimental Design
Your background or experimental plan needs to convey your plan for carrying out the experiment. This section should follow the design process as described in the Experimental Design section on page 31, and should include:
 an overview of the experiment, including any relevant equations that will be used to calculate the desired quantity
 a description of the independent variables
 a description of the dependent variables
 a description of the control variables
 a brief description of how you will calculate the desired quantity/quantities once you have performed the experiment.
Procedure
This is a detailed description of exactly what you did to set/measure the values of each of the variables. You need to include:
 A <i>labeled</i> sketch or photograph of your experimental set-up, even if the experiment is simple. The sketch will serve to answer many questions about how you set up the experiment and most of the key equipment you used.
 A list of any significant equipment that you used other than what you labeled in your sketch. (You do not need to mention generic items like pencils and paper. Basic lab safety equipment is assumed, but mention any unusual precautions that you need to take.)
 A description of how you set up the experiment, including the values of your independent variables and how you set them.
 A description of your control variables, including their values and how you are ensuring that they remain constant.
 A description of your dependent variables and how you are measuring their values. (Do not include the values of the dependent variables here—you will present those in your Data & Observations section.)
 Any significant things you did as part of the experiment besides the ones mentioned above.

Use this space for summary and/or additional notes:

Big Ideas

 Details Unit: Laborator
Data & Observations
This is a section in which you present all of your data.
For a high school lab, it is usually sufficient to present a single data table that includes the values of your independent, control, and dependent variables for each trial. However, if you have other data or observations that you recorded during the lab, they must be listed here.
You must also include estimates of the uncertainty for each measured quantity, and your calculated uncertainty for the final quantity that your experiment is intended to determine.
Analysis
The analysis section is where you interpret your data. (Note that calculated values in the table in the Data & Observations section are actually part of your analysis, even though they appear in the Data & Observations section.) Your analysis should mirror your Experimental Design section (possibly in the same order, possibly in reverse), with the goal of guiding the reader from your data to the quantity you ultimately want to calculate or determine.
Your analysis needs to include:
• A narrative description (one or more paragraphs) of the outcome of the experiment that guides the reader from your data through your calculations to the quantity you set out to determine.
• One (and only one) sample calculation for each separate equation that you used. For example, if you calculated acceleration for each of five data points, you would write down the formula, and then choose one set of data to plug ir and show how you got the answer.
 Any calculated values that did not appear in the data table in your Data & Observations section
 For some experiments, a carefully-plotted graph showing the data points you took for your dependent vs. independent variables. Note that any graphs you include in your write-up must be drawn accurately to scale, using graph paper, and using a ruler/straightedge wherever a straight line is needed. (When an accurate graph is required, you will lose points if you include a freehand sketch instead.)

Internal Laboratory Reports

		Fage. 02
Big Ideas	Details	Unit: Laboratory
	 Quantitative error analysis. In general, most quantities in a chemistry class are calculated from equations that use multi division. Therefore, you need to use relative error: 	high school plication and
	1. Determine the uncertainty of each your measurements.	
	2. Calculate the relative error for each measurement.	
	 Combine your relative errors to get the total relative err calculated value(s). 	or for your
	 Multiply the total relative error by your calculated value absolute uncertainties (±). 	s to get the
	 Sources of uncertainty: this is a list of factors <i>inherent in yo</i> limit how precise your answer can be. 	<i>ur procedure</i> that
	You need to list one source of human-derived uncertainty (e. exactly when the reaction was finished. We declared it to h nothing appeared to be changing."), and two sources of non uncertainty (e.g., "The graduated cylinder was marked in 1 r the volume was estimated to ± 0.1 mL.")	<i>g., "</i> It was unclear ave finished when <i>-human</i> nL increments, so
	Never include mistakes, especially mistakes you aren't sure you made! A statement like "We might have written down number." or "We might have done the calculations incorrect saying, "We might be stupid and you shouldn't believe anyth report." (Any "we might be stupid" statements will not cour required number of sources of uncertainty.)	the wrong the wrong tly." is really ning else in this nt toward your
	Note, however, that if a problem <i>actually occurred</i> , and if you <i>used in your calculations anyway</i> , you need to explain what happened a estimate of the effects on your results.	d that data point and give an
	Conclusion	
	Your conclusion should be worded similarly to your objective, but your final calculated result(s) and uncertainty. You do not need to of uncertainty in your conclusions unless you believe they were sig to create some doubt about your results.	this time including restate sources gnificant enough
	Your conclusion should also include 1–2 sentences describing way could be improved. These should specifically address the sources that you listed in the analysis section above.	s the experiment of uncertainty

Internal Laboratory Reports



Formal Laboratory Reports

Unit: Laboratory

Details

MA Curriculum Frameworks (2016): SP3, SP8

Mastery Objective(s): (Students will be able to...)

• Write a formal (journal article-style) laboratory report that appropriately communicates all of the necessary information.

Success Criteria:

- The report has the correct sections in the correct order.
- Each section contains the appropriate information.
- The report contains an abstract that conveys the appropriate amount of information.

Tier 2 Vocabulary: abstract

Language Objectives:

- Understand and be able to describe the sections of a formal laboratory report, and which information goes in each section.
- Write a formal laboratory report with the correct information in each section.

Notes:

A formal laboratory report serves one important purpose: to communicate the results of your experiment to other scientists outside of your laboratory or institution.

A formal report is a significant undertaking. In a research laboratory, you might submit as many as one or two articles to a scientific journal in a year. Some college professors require students to submit lab reports in journal article format.

The format of a formal journal article-style report is as follows:

Abstract

Details

Big Ideas

This is the most important part of your report. It is a (maximum) 200-word executive summary of everything about your experiment—the procedure, results, analysis, and conclusions. In most scientific journals, the abstracts are searchable via the internet, so it needs to contain enough information to enable someone to find your abstract, and after reading it, to know enough about your experiment to determine whether or not to purchase a copy of the full article (which can sometimes cost \$100 or more). It also needs to be short enough that the person doing the search won't just say "TL; DR" ("Too Long; Didn't Read") and move on to the next abstract.

Because the abstract is a complete summary, it is always best to write it last, after you have already written the rest of your report.

Introduction

Your introduction is actually an mini research paper on its own, including citations. (For a high school lab report, it should be 1–3 pages; for scientific journals, 5–10 pages is not uncommon.) Your introduction needs to describe any general background information that another scientist might not know, plus all of the background information that specifically led up to your experiment. Assume that your reader has a similar knowledge of chemistry as you, but does not know anything about this experiment. The introduction is usually the most timeconsuming part of the report to write.

Materials and Methods

This section combines both the experimental design and procedure sections of an informal lab write-up. Unlike an informal write-up, the Materials and Methods section of a formal report is written in paragraph form, in the past tense, using the passive voice, and avoiding pronouns. As with the informal write-up, a labeled photograph or drawing of your apparatus is a necessary part of this section, but you need to *also* describe the set-up in the text.

Also unlike the informal write-up, your Materials and Methods section needs to give some *explanation* of your choices of the values used for your control and independent variables.

Big Ideas	Details L	Jnit: Laboratory
	Data and Observations	
	This section is similar to the same section in the lab notebook write-u	p, except that:
	 You should present only data you actually recorded/measure section. (Calculated values are presented in the Discussion section) 	d in this ection.)
	 You need to <i>introduce</i> the data table. (This means you need to important things someone should notice in the table first, and something like "Data are shown in Table 1.") 	to describe the d then say
	Note that all figures and tables in the report need to be numbered se consecutively.	parately and
	Discussion	
	This section is similar to the Analysis section in the internal report, bu important differences.	t with some
	As with the rest of the formal report, your discussion must be in paragroup of the formal report, your discussion must be in paragroup of the formal report, your discussion your results and mean. You need to introduce and present a table with your calculate your uncertainty. After presenting the table, you should discuss the runcertainties, and sources of uncertainty in detail. If your results relate experiments, you need to discuss the relationship and include citation other experiments.	graph form. what they d values and esults, ite to other ns for those
	Your discussion needs to include each of the formulas that you used a discussion and give the results of the calculations, but you do not nee intermediate step of substituting the numbers into the equation.	as part of your d to show the
	Conclusions	
	Your conclusions are written much like in the internal write-up. You r two paragraphs. In the first, restate your findings and summarize the sources of uncertainty. In the second paragraph, list and explain impl and/or follow-up experiments that you suggest.	need at least significant rovements
	Works Cited	
	As with a research paper, you need to include a complete list of biblic for the references you cited in your introduction and/or discussion se	graphy entries ctions.
	Your ELA teachers probably require MLA-style citations; scientific pap use APA style. However, in a high school chemistry class, while it is in you know which information needs to be cited and <i>what</i> information into each citation, you may use any format you like as long as you use consistently.	ers typically nportant that needs to go it
	Use this space for summary and/or additional notes:	

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Formal Laboratory Reports

Big Ideas	Details Unit: Laboratory
	Typesetting Superscripts and Subscripts
	Because formal laboratory reports need to be typed, and because chemistry uses superscripts and subscripts extensively, it is important to know how to typeset superscripts and subscripts.
	You can make use of the following shortcuts:
	superscript: text that is raised above the line, such as the " $3+$ " in Al ³⁺ .
	In most Microsoft programs, select the text, then hold down "Ctrl" and "Shift" and press the "+" key.
	On a Macintosh, select the text, then hold down "Command" and "Control" and press the "+" key.
	subscript: text that is lowered below the line, such as the "2" in CaCl ₂ .
	In most Microsoft programs, select the text, then hold down "Ctrl" and press the "—" key.
	On a Macintosh, select the text, then hold down "Command" and "Control" and press the "-" key.
	Note that you will lose credit in your laboratory reports if you don't use superscripts and subscripts correctly. For example, you will lose credit if you type NO3– instead of NO_3^- .

Use this space for summary and/or additional notes:

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	Introduction: Math & Measurement Page	: 6
Big Ideas	Details Unit: Math & Measurem	er
	Introduction: Math & Measurement	
	Unit: Math & Measurement	
	Topics covered in this chapter:	
	The International System of Units7	0'
	Scientific Notation	7
	Using Math in Calculations8	31
	Conversions (Factor-Label Method)	8
	Dimensional Analysis	13
	Logarithms	19
	Standards addressed in this chapter:	
	Massachusetts Curriculum Frameworks & Science Practices (2016):	
	This chapter addresses the following MA science and engineering practices	:
	Practice 4: Analyzing and Interpreting Data	
	Practice 5: Using Mathematics and Computational Thinking	
	Practice 8 : Obtaining, Evaluating, and Communicating Information	

Unit: Math & Measurement

MA Curriculum Frameworks (2016): SP5

Mastery Objective(s): (Students will be able to...)

• Use and convert between metric prefixes attached to units.

Success Criteria:

Details

- Conversions between prefixes move the decimal point the correct number of places.
- Conversions between prefixes move the decimal point in the correct direction.
- The results of conversions have the correct answers with the correct units, including the prefixes.

Tier 2 Vocabulary: prefix

Language Objectives:

• Set up and solve problems relating to the concepts described in this section.

Notes:

This section is intended to be a brief review. You learned to convert between metric prefixes in elementary or middle school. You are expected to be able to fluently perform calculations that involve converting between metric prefixes.

A unit is a specifically defined measurement. Units describe both the type of measurement, and a base amount.

For example, 1 cm and 1 inch are both lengths. They are used to measure the same dimension, but the specific amounts are different. (In fact, 1 inch is exactly 2.54 cm.)

Every measurement is a number multiplied by its units. In algebra, the term "3x" means "3 times x". Similarly, the distance "75 m" means "75 times the distance 1 meter".

The number and the units are <u>both</u> necessary to describe any measurement. You *always* need to write the units. Saying that "12 is the same as 12 g" would be as ridiculous as saying "12 is the same as 12×3 ".

The International System of Units

Big Ideas Details

The International System (often called the metric system) is a set of units of measurement that is based on natural quantities (on Earth) and powers of 10.

The metric system has 7 fundamental "base" units:

Unit	Quantity	Currently Based On
meter (m)	length	the distance light travels in a specific time
kilogram (kg)	mass	the mass of the official prototype kilogram
second (s)	time	the time it takes for a particular type of radiation from a cesium-133 atom
Kelvin (K)	temperature	the temperature of the triple point of water
mole (mol)	amount of substance	the number of atoms in a specific mass of carbon-12
ampere (A)	electric current	the amount of current that produces a specific force under specific conditions
candela (cd)	intensity of light	the amount of light per unit of area at a specific distance

All other S.I. units are combinations of one or more of these seven base units.

For example:

Velocity (speed) is a change in distance over a period of time, which would have units of distance/time (m/s).

Force is a mass subjected to an acceleration. Acceleration has units of distance/time² (m/s²), and force has units of mass × acceleration. In the metric system this combination of units (kg·m/s²) is called a Newton, which means: 1 N = 1 kg·m/s²

As of 2018, Each of these base units is defined in some way that could be duplicated in a laboratory anywhere on Earth (except for the kilogram, which is defined by a physical object in a safe in France).

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	The International System of Units Page: 72
Big Ideas	Details Unit: Math & Measurement
	In May 2019, all of the above S.I. units will be defined based on specifying exact values for certain fundamental constants:
	• The Planck constant <i>h</i> is exactly $6.62607015 \times 10^{-34}$ J·s
	• The elementary charge e is exactly 1.602 176 634 × 10 ⁻¹⁹ C
	• The Boltzmann constant k is exactly $1.380649 \times 10^{-23} \text{ J} \cdot \text{K}^{-1}$
	• The Avogadro constant N_A is exactly 6.022 140 76 × 10 ²³ mol ⁻¹
	• The speed of light c is exactly 299 792 458 m·s ⁻¹
	 The ground state hyperfine splitting frequency of the caesium-133 atom Δv(¹³³Cs)_{hfs} is exactly 9 192 631 770 Hz
	• The luminous efficacy K_{cd} of monochromatic radiation of frequency 540 × 10 ¹² Hz is exactly 683 lm·W ⁻¹
	The S.I. base units are calculated from these seven definitions, after converting the derived units (joule, coulomb, hertz, lumen and watt) into the seven base units (second, meter, kilogram, ampere, kelvin, mole and candela).
 Rules for Writing S.I. Numbers and their Units The value of a quantity is written as a number followed by a space 	
--	
 The value of a quantity is written as a number followed by a space 	
(representing a multiplication sign) and a unit symbol; <i>e.g.</i> , 2.21 kg, 7.3×10^2 m ² , or 22 K. This rule explicitly includes the percent sign (10 %, not 10%) and the symbol for degrees of temperature (37 °C, not 37 °C). (However, note that angle measurements in degrees are written next to the number without a space.)	
 Units do not have a period at the end, except at the end of a sentence. 	
 A prefix is part of the unit and is attached to the beginning of a unit symbol without a space. Compound prefixes are not allowed. 	
 Symbols for derived units formed by multiplication are joined with a center dot (·) or a non-breaking space; <i>e.g.</i>, N·m or N m. 	
• Symbols for derived units formed by division are joined with a solidus (fraction line), or given as a negative exponent. <i>E.g.,</i> "meter per second" can be written m/s, m s ⁻¹ , m s ⁻¹ , or $\frac{m}{s}$.	
• The first letter of symbols for units derived from the name of a person is written in upper case; otherwise, they are written in lower case. <i>E.g.</i> , the unit of pressure is named after Blaise Pascal, so its symbol is written "Pa" (note that "Pa" is a two-letter symbol), but the symbol for mole is written "mol". However, the symbol for liter is "L" rather than "l", because a lower case "l" can be confused with the number "1".	
 A plural of a symbol must not be used; e.g., 25 kg, not 25 kgs. 	
 Units and prefixes are case-sensitive. <i>E.g.</i>, the quantities 1 mW and 1 MW represent two different quantities (milliwatt and megawatt, respectively). 	
• The symbol for the decimal marker is either a point or comma on the line. In practice, the decimal point is used in most English-speaking countries and most of Asia, and the comma in most of Latin America and in continental European countries.	
 Spaces should be used as a thousands separator (1 000 000) in contrast to commas (1,000,000) or periods (1.000.000), to reduce confusion resulting from the variation between these forms in different countries. 	
 Any line-break inside a number, inside a compound unit, or between number and unit should be avoided. 	
Is this space for summary and/or additional notas:	

The International System of Units

ts Page: 74 Unit: Math & Measurement

Big Ideas	Details

Prefixes

The metric system uses prefixes to indicate multiplying a unit by a power of ten. There are prefixes for powers of ten from 10^{-24} to 10^{24} but in chemistry, only the following four are commonly used:

• kilo (k) =
$$10^3 = 1000$$

• milli (m) = $10^{-3} = \frac{1}{1000} = 0.001$
• centi (c) = $10^{-2} = \frac{1}{100} = 0.01$
• micro (μ) = $10^{-6} = \frac{1}{1000000} = 0.00001$

Any metric prefix is allowed with any metric unit. For example, if a mole (mol) is 6.02×10^{23} objects, then a millimole (mmol) would be

 $(6.02 \times 10^{23}) \times \frac{1}{1000} = 6.02 \times 10^{20}$ objects.

An easier way to convert is to use the powers of ten that correspond with the prefixes to determine how many places to move the decimal point.

Factor		Prefix	Symbol
1 000 000 000 000 000 000 000 000	10 ²⁴	yotta	Y
1 000 000 000 000 000 000 000	10 ²¹	zeta	Z
1 000 000 000 000 000 000	10 ¹⁸	exa	E
1 000 000 000 000 000	10 ¹⁵	peta	Р
1 000 000 000 000	10 ¹²	tera	Т
1 000 000 000	10 ⁹	giga	G
1 000 000	10 ⁶	mega	М
1 000	10 ³	kilo	k
100	10 ²	hecto	h
10	10 ¹	deca	da
1	10 ⁰	—	_
0.1	10 ⁻¹	deci	d
0.01	10 ⁻²	centi	С
0.001	10 ⁻³	milli	m
0.000 001	10 ⁻⁶	micro	μ
0.000 000 001	10 ⁻⁹	nano	n
0.000 000 000 001	10 ⁻¹²	pico	р
0.000 000 000 000 001	10 ⁻¹⁵	femto	f
0.000 000 000 000 000 001	10 ⁻¹⁸	atto	а
0.000 000 000 000 000 000 001	10 ⁻²¹	zepto	z
0.000 000 000 000 000 000 000 001	10 ⁻²⁴	yocto	у

Metric Prefixes

The International System of Units

Big Ideas	Details	Unit: Math & Measurement
	Note that some of the prefixes 10 ³ . This means you can't just exponents.	skip by a factor of 10 and others skip by a factor of count the steps—you have to actually look at the
	Sample Problem:	
	15 Tm = nm	
	 You need to move the deployed places to get to 10⁻⁹, for 	ecimal point 12 places to get to 10 ⁰ , and 9 more a total of 21 places.
	 Terameters are huge, an going to have a lot more the decimal point in the 	d nanometers are much smaller. That means we're nanometers than terameters, so we have to move direction that makes the number larger (to the right).
	Therefore, we need to move the we need to multiply by 10^{21} .	e decimal point 21 places to the right, which means
	You could simply write your an you can enter it into your calcu	swer as 15×10^{21} m, and it would be correct. (And lator that way and the right thing will happen.)
	However, to be proper scientific multiplication sign between 1 a number before the \times sign gets get larger so the end result statistication 1.5×10^{22} m, which is our final a	ic notation, you need to make the part before the and 10, which means you need to make it 1.5. If the smaller, then the number after the × sign needs to ys the same. Therefore, 15×10^{21} m is the same as unswer.
	There is a popular joke based of said to have been the most bea kidnapped, the Trojan War was described as "the face that lau be the amount of beauty need	on the ancient Greek heroine Helen of Troy. She was autiful woman in the world, and when she was s fought to bring her back to Sparta. Her beauty was nched a thousand ships." Therefore a milliHelen must ed to launch one ship.

Use this space for summary and/or additional notes:

_

The International System of Units

Big Ideas	Details	Unit: Math & Measurement
	Homework Pro	oblems
	Perform the following conversions.	
	1. 2.5 m = cm	
	2. 18mL = L	
	2 68 ki -	
	5. 08 KJ = J	
	4. 6 500 mg = kg	
	5. 101 kPa = Pa	
	6 325 ms - s	
	0. 325 113 3	

Scientific Notation

Unit: Math & Measurement

MA Curriculum Frameworks (2016): SP5

Mastery Objective(s): (Students will be able to...)

• Correctly use numbers in scientific notation in mathematical problems.

Success Criteria:

Details

- Numbers are converted correctly to and from scientific notation.
- Numbers in scientific notation are correctly entered into a calculator.
- Math problems that include numbers in scientific notation are set up and solved correctly.

Language Objectives:

• Explain how numbers are represented in scientific notation, and what each part of the number represents.

Notes:

This section is intended to be a brief review. You learned to use scientific notation in elementary or middle school. You are expected to be able to fluently perform calculations that involve numbers in scientific notation, and to express the answer correctly in scientific notation when appropriate.

Scientific notation is a way of writing a very large or very small number in compact form. The value is always written as a number between 1 and 10 multiplied by a power of ten.

For example, the number 1000 would be written as 1×10^3 . The number 0.000075 would be written as 7.5×10^{-5} . The number 602 000 000 000 000 000 000 would be written as 6.02×10^{23} . The number

Scientific notation is really just math with exponents, as shown by the following examples:

$$5.6 \times 10^3 = 5.6 \times 1000 = 5600$$

$$2.17 \times 10^{-2} = 2.17 \times \frac{1}{10^{2}} = 2.17 \times \frac{1}{100} = \frac{2.17}{100} = 0.0217$$

Notice that if 10 is raised to a positive exponent means you're multiplying by a power of 10. This makes the number larger, and the decimal point moves to the right. If 10 is raised to a negative exponent, you're actually dividing by a power of 10. This makes the number smaller, and the decimal point moves to the left.

Scientific Notation

Big Ideas	Details Unit: Math & Measu	rement
	Significant figures are easy to use with scientific notation: all of the digits before	ore the
	"×" sign are significant. The power of ten after the "×" sign represents the	
	(insignificant) zeroes, which would be the rounded-off portion of the number	. In
	fact, the mathematical term for the part of the number before the "x" sign is	the
	significand.	
	Math with Scientific Notation	
	Because scientific notation is just a way of rewriting a number as a mathemat	ical
	expression, all of the rules about how exponents work apply to scientific nota	tion.
	Adding & Subtracting: adjust one or both numbers so that the power of ten is same, then add or subtract the significands.	s the
	$(3.50 \times 10^{-6}) + (2.7 \times 10^{-7}) = (3.50 \times 10^{-6}) + (0.27 \times 10^{-6})$	
	$=(3.50+0.27)\times10^{-6}=3.77\times10^{-6}$	
	<u>Multiplying & dividing</u> : multiply or divide the significands. If multiplying, add exponents. If dividing, subtract the exponents.	the
	$\frac{6.2 \times 10^8}{3.1 \times 10^{10}} = \frac{6.2}{3.1} \times 10^{8-10} = 2.0 \times 10^{-2}$	
	Exponents: raise the significand to the exponent. Multiply the exponent of the power of ten by the exponent to which the number is raised.	ıe
	$(3.00 \times 10^8)^2 = (3.00)^2 \times (10^8)^2 = 9.00 \times 10^{(8 \times 2)} = 9.00 \times 10^{16}$	
	Use this space for summary and/or additional notes:	

Using Scientific Notation on Your Calculator

Scientific calculators are designed to work with numbers in scientific notation. It's possible to can enter the number as a math problem (always use parentheses if you do this!) but math operations can introduce mistakes that are hard to catch.

Scientific calculators all have either an "EE" or "EXP" button. The entire purpose of this button is to enter numbers in scientific notation and make sure the calculator stores them properly. On Texas Instruments calculators, such as the TI-30 or TI-89, you would do the following:

What you type	What the calculator shows	What it means
6.6 EE -34	6.6E-34	6.6×10^{-34}
1.52 EE 12	1.52E12	1.52 × 10 ¹²
-4.81 EE -7	-4.81E-7	-4.81×10^{-7}

On some calculators, the scientific notation button is labeled EXP instead of EE.

Important notes:

Details

- Many high school students are afraid of the *EE* button because it is unfamiliar. If you are afraid of your *EE* button, you need to get over it and start using it anyway. However, if you insist on clinging to your phobia, you need to at least use parentheses around all numbers in scientific notation, in order to minimize the likelihood of PEMDAS errors in your calculations.
- Regardless of how you enter numbers in scientific notation into your calculator, always place parentheses around the denominator of fractions.

$$\frac{2.75 \times 10^{3}}{5.00 \times 10^{-2}} \text{ becomes } \frac{2.75 \times 10^{3}}{(5.00 \times 10^{-2})}$$

• You need to <u>write</u> answers using correct scientific notation. For example, if your calculator displays the number 1.52E12, you need to write 1.52×10^{12} (plus the appropriate unit, of course) in order to receive credit.

Use this space for summary and/or additional notes:

Big Ideas

Scient

Scientific Notation	Page: 80
Details	Unit: Math & Measurement
Homework Probl	ems
Convert each of the following between scientific and	d algebraic notation.
1. $2.65 \times 10^9 =$	
2. 387 000 000 =	
3. 1.06 × 10 ⁻⁷ =	

4. 0.000 000 065 =

3. $1.06 \times 10^{-7} =$

Solve each of the following on a calculator that can do scientific notation.

```
5. (2.8 \times 10^6)(1.4 \times 10^{-2}) =
     Answer: 3.9 \times 10^4
6. \frac{3.75 \times 10^8}{1.25 \times 10^4} =
     Answer: 3.00 \times 10^4
7. \frac{1.2 \times 10^{-3}}{5.0 \times 10^{-1}} =
      Answer: 2.4 \times 10^{-3}
```

Use this space for summary and/or additional notes:

Big Ideas

Unit: Math & Measurement

MA Curriculum Frameworks (2016): SP5

Mastery Objective(s): (Students will be able to...)

• Substitute values for variables in equations and solve them.

Success Criteria:

Details

- Values are substituted for the correct variables.
- Equations are correctly solved for the missing variable using basic algebra.
- Answers have the correct units and are rounded to the appropriate number of significant figures.

Language Objectives:

• Set up and solve word problems relating to the concepts described in this section.

Notes:

Unlike biology, chemistry is a physical science. Among other things, this means chemistry involves calculations, which means you need to be comfortable with algebraic expressions.

Variables and Units

Unlike expressions in math class, which make a clear distinction between constants (the numbers you know the value of) and the variables,

- Equations in chemistry are written as all variables, because each equation works the same way no matter which quantity (or quantities) you are looking for.
- Each of the variables is a letter that relates to the quantity that it represents. For example, volume is *V*, mass is *m*, temperature is *T*, and the number of moles of substance is *n*. In chemistry, the same quantity *always* uses the same variable.

0	Details				Unit: Math	& Measureme
	 Almost all quar to what kind of 12.5 mL, the m "grams." In the 	itities are quantity L means ' e quantity	measured the numbe "milliliters." 21.5 °C, th	and have units. ers describe. Fo ' In the quantit ne °C means "d	These un or example, i y 37.21 g, th egrees Celsi	its are your ke n the quantity ie g means us."
	 The unit is 1.62. (1.62 (1.62 mete) 	part of th what?) rs).	e quantity. You would	For example, y need to say tha	/ou can't say t your heigh	/ your height is t is 1.62 m
	 The unit te you the var is a volume equation ir letter V wit 	lls you wh iable. Fo y which u which th the qua	nich type of or example, uses the var ne letter V r antity 12.5	quantity, and t in the quantity iable "V." This represents volu mL.	he type of q 12.5 mL, the means that me, you wo	uantity tells e mL (milliliter: if you have an uld <i>replace</i> the
	 Be careful! example, the distance), the distance 	In many ne letter ' out the va	cases, the "m" next to rriable " <i>m</i> "	same letter can a number mea in an equation	be a unit or ns the unit ' means "mas	a number. Fc 'meters" (a ss".
	Quantity	Unit	Variable	Ouantity	Unit	Variable
	mass	g	m	temperature	°C, K	T
	length	m, cm	l	velocity	<u>m</u> s	v
		2				
	area	m	Α	heat	J	<i>q</i> *
	area volume	m² mL	A V	heat energy	J	q* E
	area volume number of moles	m² mL mol	A V n	heat energy pressure	J J bar, atm, kPa	q* E P
	area volume number of moles density	m² mL mol <u>g</u> mL	Α V n ρ [†]	heat energy pressure time	J J bar, atm, kPa s	q* E P t
	area volume number of moles density concentration	m ² mL mol <u>g</u> mL L	Α V n ρ ⁺ c	heat energy pressure time equilibrium constant	J J bar, atm, kPa s 	q* Е Р t К

* Notice that *q* is used for both heat and electrical charge. You need to figure out which quantity is meant from context.

⁺ Some chemistry books use the Roman letter "D" for density, but the Greek letter " ρ " ("rho") is preferred. Be careful not to confuse it with the letter "P" (pressure).

Big Ideas	Details	0	Unit: Math & Measurement
		Variable Substitu	tion
	Variable substitution and substituting tho simple version of the	n simply means taking the numb se numbers for the correspondi s is a density problem:	ers you have from the problem ng variable in an equation. A
	If you have the form	nula:	
	$\rho = \frac{m}{V}$	and you're given: $m = 12.3$	g and $V = 2.8 \text{ cm}^3$
	simply substitute 12	.3 g for m , and 2.8 cm ³ for V , giv	/ing:
		$\rho = \frac{12.3 \mathrm{g}}{2.8 \mathrm{cm}^3} = 4.4 \frac{\mathrm{g}}{\mathrm{cm}^3}$	-

Big Ideas Details

Equations

Math is a language. Like other languages, it has nouns (numbers), pronouns (variables), verbs (operations), and sentences (equations), all of which must follow certain rules of syntax and grammar.

This means that turning a word problem into an equation is translation from English to math.

Mathematical Operations

You have probably been taught translations for most of the common math operations:

word	meaning	word	meaning
and, more than		ic	_
(but not "is more than")	+	15	-
less than	_	is at least	\$
(but not "is less than")		is at least	2
of	×	is more than	>
per	÷	is at most	≤
percent	÷100	is less than	<
change in <i>x,</i> difference in <i>x</i>	Δx^*		

Suppose you were given the equation:

$$\rho = \frac{m}{V}$$

Using the table on page 82, we can see that m is mass and V is volume, which means the equation says "density is mass divided by volume". This means that if we knew that the mass of an object was 10.5 g and its volume was 23.7 mL, we could substitute those numbers into the equation to find the density:

$$\rho = \frac{m}{V} = \frac{10.5 \text{ g}}{23.7 \text{ mL}} = \frac{10.5}{23.7 \text{ mL}} = 0.443 \frac{\text{g}}{\text{mL}}$$

Note: The Greek letter Δ (delta) is attached to a variable to indicate the change in that variable. For example, ΔT represents a change in temperature. ΔT is one variable in the equation, even though it uses two symbols.

We can use the same approach no matter which variable we are looking for.

Sample Problems:

Details

Big Ideas

Q: An object has a volume of 17.7 mL and a density of $2.35 \frac{g}{mL}$. What is its mass?

A: Start with the equation and substitute:

$$\rho = \frac{m}{V}$$
$$2.35 \frac{g}{mL} = \frac{m}{17.7 \text{ mL}}$$

Now we have to do algebra. We want to get *m* by itself, which means we need to move 17.7 mL to the other side. Because it's in the denominator (on the bottom), we have to multiply both sides by it.

$$(17.7 \text{ prf})(2.35 \frac{\text{g}}{\text{rrf}}) = \frac{m}{17.7 \text{ mL}}(17.7 \text{ mL})$$

 $41.6 \text{ g} = \frac{m}{17.7 \text{ mL}}$

(Notice that the 17.7 mL cancels on the right because it's in both the numerator and the denominator. Notice also that the mL cancels on the left for the same reason, leaving g, which happens to be the correct unit. This is called "dimensional analysis," and we will study it in more depth in a future section.)

Q: An object has a mass of 44.7 g and a density of $1.68 \frac{g}{mL}$. What is its volume?

A: Again, start with the equation and substitute:

$$\rho = \frac{m}{V}$$

$$1.68 \frac{g}{mL} = \frac{44.7 \text{ g}}{V}$$

The variable we want is on the bottom. Again, following the rules of algebra, in order to get it off the bottom, we first have to multiply both sides by it to clear the fraction:

$$\mathbf{V} \cdot 1.68 \frac{g}{mL} = \frac{44.7 g}{4} \cdot \mathbf{V}$$

Then, we can solve for V in a subsequent step:

$$\frac{V \cdot 1.68 \frac{g}{mL}}{1.68 \frac{g}{mL}} = \frac{44.7 g}{1.68 \frac{g}{mL}}$$
$$\frac{V}{1.68 \frac{g}{mL}} = 26.6 \text{ mL}$$

ig ideas Unit: MAth & Measurement Note: Whenever you have to solve an equation for a quantity in the denominator, always do it in two steps: clear the fraction first, then divide. If you try to cleverly rearrange the quantities without doing this, you are almost certain to get the wrong answer! Q: Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. A: When you see a problem like this, the first thing you should do is use the units to figure out what quantities you have in the problem, and label them with the variables: V P Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. (For this problem, use $0.0821\frac{4000}{mot}$ for the gas constant.) To solve this problem, we need an equation that relates $V, n, T, and P$. This turns out to be the ideal gas law: $P V = n RT$ Now substitute the numbers in place of the variables in the equation: $P = \frac{(3.10 \text{ mol})(0.0821\frac{4000}{mot})(298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821\frac{4000}{mot})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. . Assign variables to those quantities. . Substitute the values of the variables into the equation. You should have only		
Note: Whenever you have to solve an equation for a quantity in the denominator, always do it in two steps: clear the fraction first, then divide. If you try to cleverly rearrange the quantities without doing this, you are almost certain to get the wrong answer! Q: Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. A: When you see a problem like this, the first thing you should do is use the units to figure out what quantities you have in the problem, and label them with the variables: V = n P Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. (For this problem, use $0.0821\frac{4200}{2000}$ for the gas constant.) To solve this problem, we need an equation that relates <i>V</i> , <i>n</i> , <i>T</i> , and <i>P</i> . This turns out to be the ideal gas law: P V = n RT Now substitute the numbers in place of the variables in the equation: $\frac{P}{(1.25 \text{ atm})} V = (3.10 \text{ mol}) (0.0821\frac{4200}{\text{molef}}) (298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})[(0.0821\frac{4200}{\text{molef}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!	Big Ideas	Details Unit: Math & Measuremen
Q: Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. A: When you see a problem like this, the first thing you should do is use the units to figure out what quantities you have in the problem, and label them with the variables: V = R T Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. (For this problem, use $0.0821\frac{4 \text{ state}}{R}$ for the gas constant.) To solve this problem, we need an equation that relates <i>V</i> , <i>n</i> , <i>T</i> , and <i>P</i> . This turns out to be the ideal gas law: P V = n RT Now substitute the numbers in place of the variables in the equation: $\left(1.25 \text{ atm}\right) V = (3.10 \text{ mol}) \left(0.0821\frac{4 \text{ state}}{\text{mok}}\right) (298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821\frac{4 \text{ state}}{\text{mok}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable [eft, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		Note: Whenever you have to solve an equation for a quantity in the denominator, always do it in two steps: clear the fraction first, then divide. If you try to cleverly rearrange the quantities without doing this, you are almost certain to get the wrong answer!
A: When you see a problem like this, the first thing you should do is use the units to figure out what quantities you have in the problem, and label them with the variables: V = RT Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. (For this problem, use 0.0821 finds for the gas constant.) To solve this problem, we need an equation that relates <i>V</i> , <i>n</i> , <i>T</i> , and <i>P</i> . This turns out to be the ideal gas law: P V = n RT Now substitute the numbers in place of the variables in the equation: $\frac{P}{(1.25 \text{ atm})} V = (3.10 \text{ mol}) (0.0821 \frac{\text{catm}}{\text{molk}}) (298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821 \frac{\text{catm}}{\text{molk}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		Q: Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm.
problem, use $0.821 \frac{1.38m}{R}$ for the gas constant.) To solve this problem, we need an equation that relates <i>V</i> , <i>n</i> , <i>T</i> , and <i>P</i> . This turns out to be the ideal gas law: PV = n RT Now substitute the numbers in place of the variables in the equation: $\begin{pmatrix} P \\ (1.25 \text{ atm}) V = (3.10 \text{ mol}) (0.0821 \frac{f}{\text{molk}}) (298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821 \frac{f}{\text{molk}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		 A: When you see a problem like this, the first thing you should do is use the units to figure out what quantities you have in the problem, and label them with thei variables: Find the volume taken up by 3.10 mol of a gas at 298 K and 1.25 atm. (For this
To solve this problem, we need an equation that relates <i>V</i> , <i>n</i> , <i>T</i> , and <i>P</i> . This turns out to be the ideal gas law: P V = n RT Now substitute the numbers in place of the variables in the equation: $\begin{pmatrix} P & V = n R \\ (1.25 \text{ atm}) V = (3.10 \text{ mol}) (0.0821 \frac{(\text{stm})}{\text{molk}}) (298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821 \frac{(\text{stm})}{\text{molk}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		problem, use $0.0821 \frac{\ell \text{ atm}}{R}$ for the gas constant.)
P V = n RT Now substitute the numbers in place of the variables in the equation: $\begin{pmatrix} P & V = n \\ (1.25 \text{ atm}) & V = (3.10 \text{ mol}) & (0.0821 \frac{ratm}{\text{molk}}) & (298 \text{ K}) \end{bmatrix}$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821 \frac{ratm}{\text{molk}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		To solve this problem, we need an equation that relates <i>V</i> , <i>n</i> , <i>T</i> , and <i>P</i> . This turns out to be the ideal gas law:
Now substitute the numbers in place of the variables in the equation: $\begin{pmatrix} P \\ (1.25 \text{ atm}) \\ V = (3.10 \text{ mol}) \\ (0.0821 \frac{fatm}{molk}) \\ (298 \text{ K}) \end{bmatrix}$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821 \frac{fatm}{molk})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), and what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		P V = n RT
$P = V = n (1.25 \text{ atm}) V = (3.10 \text{ mol}) (0.0821 \frac{\ell \text{ atm}}{\text{molk}}) (298 \text{ K})$ Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. $V = \frac{(3.10 \text{ mol})(0.0821 \frac{\ell \text{ atm}}{\text{mok}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ell$ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), and what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units!		Now substitute the numbers in place of the variables in the equation:
 Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer. V = (3.10 mol)(0.0821 (298 K))/(1.25 atm) = 60.7 l The Problem-Solving Process Identify the quantities in the problem, using the units. Assign variables to those quantities. Make a list of all of your variables (including the one you're looking for), anwhat they're equal to. Write down an equation that relates all of those variables. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		$P = N = n = R = T$ (1.25 atm) V = (3.10 mol) (0.0821 $\frac{\ell \cdot atm}{mol \cdot K}$) (298 K)
 V = (3.10 mol)(0.0821 (-atm molt M)(298 K))/(1.25 atm) = 60.7 ℓ The Problem-Solving Process 1. Identify the quantities in the problem, using the units. 2. Assign variables to those quantities. 3. Make a list of all of your variables (including the one you're looking for), an what they're equal to. 4. Write down an equation that relates all of those variables. 5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. 6. Solve the equation, using algebra. 7. Don't forget to round your answer correctly and include the units! 		Then solve, using algebra. This means we need to divide both sides by 1.25 atm to get the answer.
 The Problem-Solving Process Identify the quantities in the problem, using the units. Assign variables to those quantities. Make a list of all of your variables (including the one you're looking for), an what they're equal to. Write down an equation that relates all of those variables. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		$V = \frac{(3.10 \text{ mol})(0.0821 \frac{\ell \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})}{1.25 \text{ atm}} = 60.7 \ \ell$
 Identify the quantities in the problem, using the units. Assign variables to those quantities. Make a list of all of your variables (including the one you're looking for), an what they're equal to. Write down an equation that relates all of those variables. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		The Problem-Solving Process
 Assign variables to those quantities. Make a list of all of your variables (including the one you're looking for), an what they're equal to. Write down an equation that relates all of those variables. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		1. Identify the quantities in the problem, using the units.
 Make a list of all of your variables (including the one you're looking for), an what they're equal to. Write down an equation that relates all of those variables. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		2. Assign variables to those quantities.
 Write down an equation that relates all of those variables. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		 Make a list of all of your variables (including the one you're looking for), any what they're equal to.
 Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for. Solve the equation, using algebra. Don't forget to round your answer correctly and include the units! 		4. Write down an equation that relates all of those variables.
6. Solve the equation, using algebra.7. Don't forget to round your answer correctly and include the units!		5. Substitute the values of the variables into the equation. You should have only one variable left, which is the one you're looking for.
7. Don't forget to round your answer correctly and include the units!		6. Solve the equation, using algebra.
		7. Don't forget to round your answer correctly and include the units!
		ose and space for summary and/or additional notes.
ose this space for summary and/or additional notes.		

Big Ideas	Details	Unit: Math & Measurement			
	Home	Homework Problems			
	For these problems, use the table of which quantities represent which va equation given. <i>You do not have to s</i>	units and variables on page 82 to determine riables. Then substitute the variables into the solve the equations.			
	1. 375 J of heat is added to a 75 capacity of $C = 0.450 \frac{1}{g^{-C}}$. Wh $q = mC\Delta T$	g block of metal that has a specific heat at is the temperature change of the metal?			
	Answer: (375 J) = (75 g)(0.450	$\frac{1}{g^{\circ}C})\Delta T$			
	2. A rock has a density of $6.4 \frac{g}{cm^3}$ $\rho = \frac{m}{V}$	and a mass of 1 500 g. What is its volume.			
	 2.5 mol of an ideal gas has a p The gas constant is 0.081 Libar mol-K 	pressure of 1.5 bar and a temperature of 325 K. . What is its volume?			

Unit: Math & Measurement

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Use algebra and units to create a strategy for problem-solving.

Success Criteria:

Details

- Conversion factors are arranged so that the numerator and denominator are equal.
- Units in conversion factors are arranged so they cancel unwanted units and provide desired units.
- Answers are correct with the correct units.

Tier 2 Vocabulary: unit, convert, conversion

Language Objectives:

• Understand and explain that a conversion factor is two quantities (including their units) that are equal.

Notes:

A conversion is based on the idea that you can express the same quantity using different numbers and units.

For example, Mr. Bigler is 5 feet 4 inches tall. We could express this as 64 inches, $5\frac{1}{3}$ feet, $1.\overline{7}$ yards, 0.001 mile, 163 cm, 1.63 m, or 5.3×10^{-13} parsecs.^{*}

The process of getting from one of these numbers to another is called a unit conversion.

Conversions are based on two strategies:

- 1. Canceling units you don't want and replacing them with units you do want.
- 2. Repeatedly multiplying by fractions that equal 1 (*i.e.*, the numerator equals the denominator), so the actual quantity doesn't change.

^{*} A parsec is a distance of about 3.26 light years, or about 3×10^{13} km.

Conversions (Factor-Label Method)

Big Ideas Details

To show how this works, consider the following math problem:

$$\frac{1}{2} \times \frac{2}{3} \times \frac{3}{4} \times \frac{4}{5} \times \dots \times \frac{99}{100} = ?$$

The answer is $\frac{1}{100}$, because everything else cancels:
$$\frac{1}{2} \times \frac{2}{3} \times \frac{3}{4} \times \frac{4}{5} \times \dots \times \frac{99}{100} = \frac{1}{100}$$

As you may know from algebra, this also works with numbers and variables:

$$\left(\frac{4y}{1}\right)\left(\frac{3x}{y}\right)\left(\frac{4y}{2x}\right)\left(\frac{5z}{2y}\right) = \frac{\cancel{4}\cdot 3\cdot 5\cdot z}{\cancel{2}\cdot \cancel{2}} = 15z$$

Units work just like variables, so the algebra that you can do with a variable also works with a unit:

$$\frac{2 \text{ yd.}}{1} \times \frac{3 \text{ ft.}}{1 \text{ yd.}} \times \frac{12 \text{ in.}}{1 \text{ ft.}} = \frac{2 \times 3 \times 12 \text{ in.}}{1} = 72 \text{ in.}$$

Notice also that each time we multiplied by a fraction, the numerator was equal to the denominator. (3 ft. = 1 yd. and 12 in. = 1 ft.) This means we were multiplying by 1 each time. That's why 72 in. *must be* the <u>same</u> distance as 2 yd. We converted by multiplying:

(The "1"s are in quotes because the fractions derived from the conversion factors are all equal to one, even if they don't look like it.)

Some chemistry teachers prefer to use a table with lines to keep the conversion factors neat. The following are two equivalent ways to represent the same calculation. Note that conversion factors (fractions that equal 1) are in vertical columns:

$$\frac{2 \text{ yd.}}{1} = \frac{3 \text{ ft.}}{1 \text{ yd.}} = \frac{12 \text{ in.}}{1 \text{ in.}} = \frac{2 \times 3 \times 12 \text{ in.}}{1 \times 1} = 72 \text{ in.}$$

$$\frac{2 \text{ yd.}}{1 \text{ yd.}} \times \frac{3 \text{ ft.}}{1 \text{ yd.}} \times \frac{12 \text{ in.}}{1 \text{ ft.}} = \frac{2 \times 3 \times 12 \text{ in.}}{1} = 72 \text{ in.}$$

Big Ideas	Details	Unit: Math & Measurement		
	To conv	ert a quantity from one unit to another:		
	1.	Write down the number <u>and units</u> that you're starting with		
	2.	Find a conversion factor that contains the unit you want to get rid of. Turn the conversion into a fraction and arrange it to cancel the unit you want to get rid of. (If the unit you want to cancel is in the numerator, the same unit needs to be in the denominator in the fraction.) Repeat steps 2 & 3 until you end up with the unit you want.		
	3.			
	4.			
	5.	After canceling units, multiply and divide the numbers in the numerator & denominator and simplify the expression.		
	Working	g Example:		
	Wha	at is the mass (in grams) of 2.75 moles of sodium chloride (NaCl)?		
	Con	version factor for NaCl:		
		1 mol = 58.44 g		
	Solution	1:		
		1. We are starting with 2.75 moles (2.75 mol) of chlorine. This means we		
		need to write 2.75 mol in fraction form, as $\frac{2.75 \text{ mol}}{1}$.		
		2. We want to multiply $\frac{2.75 \text{ mol}}{1} \times 1$ (so we don't change the actual amount).		
		This will become $\frac{2.75 \text{ mol}}{1} \times$ and the unknown fraction needs		
		to equal 1.		
		3. We know we need to cancel moles, so moles will end up on the bottom of the next fraction. This gives us:		
		$\frac{2.75 \text{ mol}}{1} \times \frac{1}{\text{ mol}}$		

Conversions (Factor-Label Method)

Big Ideas	Details	Unit: Math & Measurement
	4.	Our conversion factor was: 1 mole NaCl = 58.44 g NaCl which means our conversion factor is:
		$\frac{1 \text{mol}}{58.44 \text{g}} = \frac{58.44 \text{g}}{1 \text{mol}} = 1$
		We need moles on the bottom, which means we need to use the second fraction. Multiplying our original 2.75 mol by this fraction gives us:
		$\frac{2.75 \text{ met}}{1} \times \frac{58.44 \text{ g}}{1 \text{ mol}}$
	5.	Now, the only unit left is the one we want (grams), so we're ready to solve the problem:
		$\frac{2.75\text{prof}}{1} \times \frac{58.44\text{g}}{1\text{prof}} = \frac{2.75\times58.44\text{g}}{1\times1} = \frac{160.71\text{g}}{1} = 160.71\text{g}$
	6.	Because we had only 3 significant figures in the original number, we need to round our answer to 3 "sig figs". This gives us our final, rounded-off answer of 161 g.

Conversions (Factor-Label Method)

Unit: Math & Measurement **Big Ideas** Details **Homework Problems** Perform each of the following conversions. 1. 23.6 cm = _____ m 7. 64 inches = _____ cm 2. 15.9 L = _____ mL 8. 183 pounds = _____ kg 9. $65\frac{\text{miles}}{\text{hour}} = \underline{\qquad \frac{\text{m}}{\text{s}}}$ 3. 0.89 km = _____ mm 4. $7.31 \times 10^{24} \text{ mmol} = ____ \text{mol}$ 10. $13.2 \frac{g}{cm^3} = ____ \frac{\text{pounds}}{\text{foot}^3}$ 11. $3.65 \frac{\text{dollars}}{\text{gallon}} = \underline{\qquad} \frac{\text{cents}}{\text{L}}$ 5. 15.0 gallons = _____ L 12. $32 \frac{\text{miles}}{\text{gallon}} = \underline{\qquad} \frac{\text{km}}{\text{L}}$ 6. 3.65 miles = _____ km **Conversion Factors** 1 gallon = 3.785 L 1 mile = 1.61 km $(1 \text{ inch})^3 \equiv (2.54 \text{ cm})^3$ I hour $\equiv 60 \text{ min}$. 1 pound = 454 g 1 inch \equiv 2.54 cm (1 foot)³ \equiv (12 inch)³ 1 min. \equiv 60 s $1 \text{ foot} \equiv 12 \text{ inches}$

Use this space for summary and/or additional notes:

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Unit: Math & Measurement

MA Curriculum Frameworks (2016): SP5

Mastery Objective(s): (Students will be able to...)

- Use compound units to infer conversion factors.
- Infer equations from compound units.

Success Criteria:

Details

- Equations relate quantities in the same way that units do.
- Word problems involving conversions that use inferred quantities are correct, including the correct units, and are rounded to the appropriate number of significant figures.

Tier 2 Vocabulary: dimension, unit, conversion

Language Objectives:

• Explain the process of using a compound unit as a conversion factor using appropriate vocabulary.

Notes:

dimensions: the units that something is measured in.

<u>dimensional analysis</u>: using the relationship between units (dimensions) and equations in order to analyze (and solve) problems. This can involve either:

- Using units to predict the relationships between quantities (and sometimes the equations that relate them)
- Using an equation to determine what the dimensions (units) should be.

Remember that units are like variables. They can be multiplied and divided by their coëfficients, and by each other. If you divide meters (a unit of distance) by seconds (a unit of time) you end up with the units $\frac{m}{s}$ (a unit of velocity^{*}). Because this works for one set of units, it must also work for other units that measure the same dimensions. Because we divided the units of distance by the units of time and got units of velocity, this means we can divide <u>any</u> distance (regardless of the units) by <u>any</u> time (regardless of the units) to get velocity.

^{*} There is a subtle difference between velocity and speed, which you will study if you take physics.

Big Ideas	Details	Unit: Math & Measurement	
	For example, density measures how heavy something is for its size. The density of iron is $7.87 \frac{g}{cm^3}$. Because grams measure mass and cm ³ measure volume, it must be		
	true that:		
	density units $=\frac{\text{mass units}}{\text{volume units}}$ which means	density = $\frac{\text{mass}}{\text{volume}}$	
	This must mean that <u>any</u> mass unit divided by <u>any</u> volum (which is, in fact, true).	e unit gives a density unit	
	Some of the units that density can be measured in (mass units) include:	units divided by volume	
	$\frac{g}{mL} \frac{g}{L} \frac{g}{cm^3} \frac{kg}{L} \frac{kg}{m^3} \frac{tonne}{m^3}$	$\frac{\text{lb.}}{\text{ft.}^3}$	
	You can use dimensional analysis in the same manner to solve many problems that include units of measure. In fa has led to new scientific discoveries. Scientists can some relationship between two quantities based on the units, that relationship in the laboratory!	help you figure out how to act, dimensional analysis etimes determine a and are then able to prove	
	However, there are pitfalls. For example, mechanical we same units $(N \cdot m)$, but they describe different kinds of quesce used interchangeably. (In fact, in a physics problem invo a wrench, 36 N \cdot m of torque ended up doing 19 N \cdot m of v	ork and torque have the uantities and cannot be Iving tightening a bolt with work on the bolt!)	
	Also, dimensional analysis cannot predict constants that equation. For example, the unit for energy is a joule, wh	might appear in an ich equals one $\frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}$.	
	Because mass is measured in kg and velocity in $\frac{m}{s}$, this s	uggests that the equation	
	for kinetic energy should be mv^2 . However, the equation	is actually $\frac{1}{2}mv^2$. The	
	factor of $\frac{1}{2}$ is a necessary part of the equation, but cannot	ot be discovered by	
	dimensional analysis.		

Making Compound Units Into Conversion Factors

Any time you have a number with a compound unit, you can make a "temporary" conversion factor out of it. For example, if a car is driving $60 \frac{\text{mi.}}{\text{hr}}$, then:

$$60 \frac{\text{mi.}}{\text{hr.}} = 60 \times \frac{\text{mi.}}{\text{hr.}} = \frac{7.87}{1} \times \frac{\text{mi.}}{\text{hr.}} = \frac{60 \text{ mi.}}{1 \text{ hr.}}$$

In other words, the quantity $60 \frac{\text{mi.}}{\text{hr.}}$ gave us the conversion factor 60 mi. = 1 hr.(Notice that the coëfficient of 60 ended up with miles, which is the unit that was on top.) This means that for this problem—as long as the car keeps going the same speed—60 miles takes 1 hour. That means you can use this conversion factor *in this problem* to convert miles to hours, or hours to miles.

Examples:

• The density of iron (Fe) is $7.87 \frac{g}{cm^3}$. What is the volume of 15.0 g of iron?

To solve this problem, we recognize that the compound unit $7.87 \frac{g}{cm^3}$ can be written as follows:

$$7.87 \frac{g}{cm^3} = 7.87 \times \frac{g}{cm^3} = \frac{7.87}{1} \times \frac{g}{cm^3} = \frac{7.87 g}{1 cm^3}$$

This means 7.87 g = 1 cm³. We can use this as the conversion factor between grams of iron and cubic centimeters of iron. We will use in this problem to convert grams of iron into cm³ of iron:

$$\frac{15.0 \text{ g}}{1} \times \frac{1 \text{ cm}^3}{7.87 \text{ g}} = \frac{15.0 \text{ cm}^3}{7.87} = 1.91 \text{ cm}^3$$

Use this space for summary and/or additional notes:

Big Ideas

Big Ideas	Details	Unit: Math & Measurement
	• The concentration of sodium chloride (Na	Cl) in sea water is about $0.48 \frac{\text{mol}}{\text{L}}$.
	How many moles of NaCl are in 55,000 L o	of sea water?
	To solve this problem, we recognize that t	the compound unit 0.48 $\frac{mol}{L}$ is going
	to be the conversion factor 0.48 mol = 1 L factor that we will use in this problem to o of NaCl:	. This is the temporary conversion convert liters of sea water into moles
	$\frac{55000\textit{k}}{1} \times \frac{0.48\text{mol}}{1\textit{k}} = 55000 \times 0.48$	mol = 26 400 mol = 26 000 mol
	(Note that we had to round off the final a	nswer to 2 significant figures.)
	• The molar mass of NaOH is $40.00 \frac{g}{mol}$. W	hat is the mass of 2.85 mol of NaOH?
	To solve this problem, we recognize that t	the compound unit $40.00 \frac{g}{mol}$ is going
	to be the conversion factor 40.00 g = 1 me factor that we will use in this problem to e	ol. This is the temporary conversion convert moles of NaOH into grams:
	$\frac{2.85\text{møl}}{1} \times \frac{40.00\text{g}}{1\text{møl}} = 2.8$	$35 \times 40.00 \text{ g} = 114 \text{ g}$

Big Ideas	Details	Unit: Math & Measurement
		Homework Problems
	1. An obj what is	ect has a density of $3.65 \frac{g}{cm^3}$. If the volume of the object is 12.5 cm ³ , s its mass?
	Answe 2. A liquio salt are	r: 45.6 g d solution has a salt concentration of 2.5 mol/l . How many moles of e in 0.50 & of the solution?
	Answe 3. A car is this ca	r: 1.3 mol s travelling at a speed of 65 mi./hr. How many hours would it take for r to travel 250 mi.?
	Answe 4. Suppos years . childre of the	r: 3.8 hr se the average temperature of the Earth is rising at a rate of 2.0°C/100 When Mr. Bigler gives this same homework assignment to one of your en 25 years from now, how much warmer will the average temperature Earth be?
	Answe	r: 0.5 °C

Big Ideas	Details Unit: Math & Measuremen
	 If a gas at "standard temperature and pressure" has a molar volume of 22.7 L/mol , how many moles of this gas would there be in a 5.5 L balloon?
	Answer: 0.24 mol
	6. Suppose you have a job at which you earn \$8.00 per hour (which you can write as 8.00 \$/hr). How many hours would you have to work to have enough money for a \$1200 car?
	Answer: 150 h
	7. The continent of South America is drifting away from Africa at a rate of about $2.5 \frac{\text{cm}}{\text{vear}}$. If South America was once touching Africa, and the speed of the
	plates was constant, how many years did it take for South America to get to its present location, which is about 5000 km away from Africa?
	(Hint: don't forget that you will need to convert km to mm.)
	Answer: ?

Logarithms

Unit: Math & Measurement

MA Curriculum Frameworks (2016): SP5

Mastery Objective(s): (Students will be able to...)

• Use logarithms to solve for a variable in an exponent.

Success Criteria:

Equations use logarithms to turn equations of the form a^x = b into equations of the form xlog(a) = log(b).

Tier 2 Vocabulary: logarithm

Language Objectives:

• Explain what the logarithm function is used for.

Notes:

Details

The logarithm may well be the least well-understood function encountered in high school mathematics. In high school chemistry, logarithms are used for the pH function for measuring the strength of acids & bases (which we will cover at the end of the year).

The simplest way to understand logarithms is to start with the base ten logarithm. You can think of the (base ten) logarithm of a number as the number of zeroes after a number.

х	X		
100 000	10 ⁵	5	
10 000	10 ⁴	4	
1 000	10 ³	3	
100	10 ²	2	
10	10 ¹	1	
1	10 ⁰	0	
0.1	10 ⁻¹	-1	
0.01	10 ⁻²	-2	
0.001	10-3	-3	
0.000 1	10 ⁻⁴	-4	
0.000 01	10 ⁻⁵	-5	

As you can see from the above table, the logarithm of a number turns a set of numbers that vary exponentially (powers of ten) into a set that vary linearly.

Logarithms

Details		•	Unit:	Math & Measurement	
You can get a vi	sual sense of t	he logarithr	n function from the l	ogarithmic number	
line below:					
3	5 7 9	30) 50 70 90 	300 500 700 900	
1 2	4 6 8 10	20	40 60 80 100	200 400 600 800 1000	
Notice that the	<i>distance</i> from	1 to 10 is th	e same as the distan	ce from 10 to 100 and	
from 100 to 100	0. In fact, the	e relative dis	stance to every numb	per on this number	
line is the logari	thm of the nu	mber.			
			distance from begin	nning	
	X	$\log_{10}(x)$	of number line	2	
	10 ⁰	0	0		
	10 ^{0.5} ≈ 3.16	0.5	½ cycle		
	$10^{1} = 10$	1	1 cycle		
	$10^2 = 100$ $10^3 = 1000$	2	2 cycles		
	10 = 1000	5	5 Cycles		
numbers get lar The most useful into the linear p	ger. mathematica art of the equ log ₁	l property o ation: ₀ (10 ³)=3lo	f logarithms is that tl g ₁₀ (10) = (3)(1) = 3	hey move an exponent	
In fact, the logarithm function works the same way for any base, not just 10:					
	lo	$g_2(2^7) = 7 \log_2(2^7)$	$g_2(2) = (7)(1) = 7$		
(In this case, the equation is:	e word "base"	means the l	base of the exponent	t.) The general	
$\log_x(a^b) = b \log_x(a)$					
This is a powerful tool in solving for the exponent in an equation. This is, in fact, precisely the purpose of using logarithms in most mathematical equations.					

		Logarithm	S	Page: 101
Big Ideas	Details		Unit: Math &	Measurement
	Sample problem	n:		
	Q: Solve $3^x = 15$ for	or x.		
	A: Take the logarit base implies th	thm of both sides. (No at the base is 10.)	ote that writing "log" without	t supplying a
	$log(3^{x}) = log(15)$ $x log(3) = log(15)$ $(x)(0.477) = 1.1$ $x = \frac{1.176}{0.477} = 2.4$	i) 5) 76 65		
	This is the corre	ect answer, because 3	^{2.465} = 15	
	A powerful tool tha logarithmic graph p you plot an expone logarithmic ("semi- graph paper that ha axis but not the oth The graph at the rig Notice where the fo the graph:	at follows from this is up opper to solve equation ential function on semi- log") graph paper (me as a logarithmic scale oper), you get a straight ght is the function $y = 3$ ollowing points appear Main Range 0 1 1 2 2 4 1 2 2 4 1 3 8 1 4 1 6 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1	Ising 50 ns. If 40 aning 30 on one line. 20 2 ^x . ron 10 9 8 7 6 5 4 2 4 4 4 4 4 4 4 4 4 4 4 4 4	
	Notice also that you intermediate value the graph shows th	5 32 u can use the graph to s. For example, at x = hat y = 6.06.	find 2.6, 10012	3 4 5

Details	Unit: Matter	
Introduction: Matter		
Unit: Matter		
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Standards addressed in this chapter:		
Massachusetts Curriculum Frameworks & Science Practices (2016):		
HS-PS1-11(MA): Design strategies to identify and separate the com a mixture based on relevant chemical and physical properties.	ponents of	
HS-PS2-8(MA) : Use kinetic molecular theory to compare the streng electrostatic forces and the prevalence of interactions that occur molecules in solids, liquids, and gases. Use the combined gas law determine changes in pressure, volume, and temperature in gase	ths of between to es.	
	Details Introduction: Matter Unit: Matter Topics covered in this chapter: States of Matter Properties of Matter Phase Diagrams Phases & Phase Changes Conservation of Mass, Energy and Charge Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-P51-11(MA) : Design strategies to identify and separate the com a mixture based on relevant chemical and physical properties. HS-P52-8(MA) : Use kinetic molecular theory to compare the streng electrostatic forces and the prevalence of interactions that occur molecules in solids, liquids, and gases. Use the combined gas law determine changes in pressure, volume, and temperature in gase	

States of Matter

Unit: Matter

Details

MA Curriculum Frameworks (2016): HS-PS1-11(MA)

Mastery Objective(s): (Students will be able to...)

- Define & describe states of matter and transitions between states.
- Classify matter according to its physical state.

Success Criteria:

- States of matter and transitions are described using correct vocabulary.
- States of matter are correctly identified based on observable properties.

Tier 2 Vocabulary: state, matter

Language Objectives:

- Explain the three common states of matter (solid, liquid and gas) and the properties of each.
- Explain the difference between a physical and chemical change.

Notes:

matter: anything that has mass and takes up space (has volume).

Examples of matter: anything you can touch or feel—solids, liquids, and gases.

Examples of things that are <u>not</u> matter: forms of energy such as light, microwaves, radio waves, etc.

state of matter: the physical form the matter is in (solid, liquid, gas, or plasma)

- <u>solid</u>: a state of matter in which the molecules are bonded (attached) to each other. Molecules in a solid move back and forth, but cannot break free from the other molecules. Solids have a definite shape and a definite volume.
- <u>liquid</u>: a state of matter in which bonds between the molecules are continuously breaking and forming. Molecules in a liquid are free to move, but are attracted to nearby liquid molecules. Liquids have a definite volume, but not a definite shape. (Liquids take on the shape of their containers.)
- gas: a state of matter in which the molecules are not bonded to one another. Molecules in a gas are free to move anywhere within the confines of their container. Gases have neither a definite shape nor a definite volume. (Gases expand to fill their containers.)

States of Matter

Big Ideas	Details		Unit: Matter
	<u>plasma</u> : a state i onto all of th and can ofter	n which the molecules I eir electrons. Charge is n be seen as blue streak	nave so much energy that they cannot hold continuously flowing through the plasma is (such as in a plasma globe).
	Note that the distinctions between the phases can be subtle. For example, ketchup has a definite shape unless you wait for a long time, but it eventually takes on the shape of its container. As it turns out, ketchup is a liquid with a high viscosity (meaning that it resists flowing). Glass flows very slowly (windows that are centuries old are thicker at the bottom than at the top), but this is because of the movement of solid particles. (This is analogous to small pebbles settling to the bottom of a bucket of rocks.) Glass is therefore an amorphous (irregularly-shaped) solid, not a viscous liquid.		
	melting: the tran	sition from a solid to a	liquid.
	freezing: the trai	nsition from a liquid to a	a solid.
	boiling: the trans	sition from a liquid to a	gas.
	<u>condensation</u> (or	condensing): the trans	ition from a gas to a liquid.
	sublimation (or s	ubliming): the transitio	n from a solid directly to a gas.
	deposition (or depositing): the transition from a gas directly to a solid.		
	Some properties between them ar	of solids, liquids and ga e summarized in the ph	ses and the processes of converting nase diagram below:
	pressure	Solid definite shape, definite volume molecules rigidly bonded together freezing m deposition sublimation Gas	Liquid indefinite shape, definite volume bonds continuously breaking & re-forming elting condensation boiling indefinite shape & volume molecules not bonded together and able to move freely
		tem	perature

Unit: Matter

Details

MA Curriculum Frameworks (2016): HS-PS1-11(MA)

Mastery Objective(s): (Students will be able to...)

- Describe the properties of mixtures and pure substances.
- Classify substances as heterogeneous or homogeneous mixtures, compounds, or elements according based on information about those substances.
- Identify methods of separating mixtures based on differences in properties.

Success Criteria:

- Substances are correctly identified as mixtures (heterogeneous or homogeneous), compounds or elements.
- Suitable methods for separating mixtures are chosen based on differences in chemical or physical properties.

Tier 2 Vocabulary: physical, chemical, property, mixture, compound, element **Language Objectives:**

• Demonstrate understanding of the key terms "homogeneous mixture," "heterogeneous mixture," "compound," and "element."

Notes:

<u>physical properties</u>: characteristics of the substance that can be measured or observed without changing the identity of the substance. *E.g.,* boiling point, freezing point, density, size, shape, color, *etc.*

<u>chemical properties</u>: characteristics of the substance having to do with how the atoms and molecules that make up substance can be combined with or changed into other substances. These properties can only be measured through changes to the identity of the substance. *E.g.*, chemical reactivity, flammability.

<u>physical change</u>: any change that alters the physical properties of the substance, such as freezing, boiling, tearing, crushing, *etc.*

<u>chemical change</u>: any change that alters the chemical properties (identity) of the substance, such as burning, cooking, rusting, decaying, *etc*.

Properties of Matter

Big Ideas	Details	Unit: Matter
	Note that the difference between a physical change and a chemical chan subtle. For example, if you have a solution of sugar dissolved in water a the water evaporate, this would be a <i>physical</i> change because the sugar molecules are each unchanged by the process.	nge can be nd you let and water
	However, if you have a solution of salt dissolved in water and you evapor water, this would be a <i>chemical</i> change, because when salt dissolves in w ionic bonds between the sodium and chloride ions break and the ions re separate while the salt is in solution. When you evaporate the water, io form between the sodium and chloride ions, which creates the ionic con sodium chloride.	rate the water, the main nic bonds npound
	There is no way to see this difference. This means you need to understa bonds and the processes of forming and breaking them in order to be ab whether a change is physical or chemical!	nd chemical le to decide
	Mixtures vs. Pure Substances	
	<u>mixture</u> : two or more different substances sharing the same space or vo Mixtures can be separated based on differences in physical properties	olume. es.
	Mixtures can be:	
	<u>homogeneous</u> : every sample of the mixture is the same, no matter we the mixture it's taken from. (<i>homo</i> = same) <i>E.g.</i> , salt water. Gate	vhat part of orade
	<u>heterogeneous</u> : samples taken from different parts of the mixture m different. (<i>hetero</i> = different) <i>E.g.,</i> chocolate chip cookies, oran	nay be ge juice.
	pure substance: a pure substance is a substance that cannot be separate broken down by any physical change. A pure substances can be a:	ed or
	<u>compound</u> : a substance made out of different kinds of atoms that a chemically bonded together. Compounds can be broken down t chemical changes. Anything that can be described by a chemica a compound. <i>E.g.</i> , H ₂ O, C ₆ H ₁₂ O ₆ , NaCl (table salt), C ₃ H ₈ O (rubbin	re hrough I formula is g alcohol).
	<u>element</u> : a substance made out of only one kind of atom. Elements broken down through chemical changes. The periodic table lists known elements according to their properties, which means any on the periodic table is an element. <i>E.g.</i> , iron, gold, oxygen, alur	cannot be all of the substance ninum.

Properties of Matter

Big Ideas	Details Unit: Matter
	Separating Mixtures
	Mixtures can be separated based on differences in the physical properties of the different substances that make up the mixture. Some processes used for separating mixtures include:
	<u>filtration</u> : separating substances by size—larger ones are trapped on the filter and smaller ones can pass through.
	distillation: separating substances that have different boiling points by heating to a temperature at which one boils and the other does not.
	evaporation: evaporating or boiling off water (or another solvent) to leave behind a solid.
	<u>crystallization</u> : separating substances that have different freezing points by letting one form a solid (freeze), but not the other.
	<u>centrifugation</u> : separating substances according to their densities by spinning them at high speeds.
	<u>chromatography</u> : separating substances by how quickly or slowly they move through another substance.
	Homework Problems
	For each of the following changes, state whether the change is chemical or physical and explain how you know.
	1. water boiling
	2. iron rusting
	3. cooking an egg
	4. breaking glass
	5. tearing up a piece of paper
	6. burning a piece of paper
	Use this space for summary and/or additional notes:
Properties of Matter

Big Ideas	Details	Unit: Matter
	7. making crushed ice in a blender	
	8. garbage turning into compost	
	9. leaves changing color	
	Classify the each of the following types of matt homogeneous mixture, compound, or element	er as a heterogeneous mixture,
	10. pure water (H_2O)	16. carbonated soda
	11. helium	17. ice water (pure H ₂ O, but both liquid and solid)
	12. chocolate-chip cookies	18. aluminum
	13. salt water	19. chicken noodle soup
	14. orange juice	20. chicken broth
	15. 14-karat gold (note: pure gold is 24K)	21. glucose (C ₆ H ₁₂ O ₆)
	22. How would you separate a mixture of s	sugar, sand, and hollow plastic beads?

Unit: Matter

Details

Big Ideas

MA Curriculum Frameworks (2016): HS-PS1-11(MA)

Mastery Objective(s): (Students will be able to...)

- Identify the phase of a substance at any combination of temperature and pressure.
- Determine the melting and boiling points of a substance any pressure.

Success Criteria:

- Phases are correctly identified as solid, liquid, gas, supercritical fluid, *etc.*, in accordance with the temperature and pressure indicated on the phase diagram.
- Melting and boiling point temperatures are correctly identified for a substance from its phase diagram for a given pressure.
- The effects of a pressure or temperature change (*e.g.,* substance would melt, sublime, *etc.*) are correctly explained based on the phase diagram.

Tier 2 Vocabulary: phase, curve, fusion, solid, liquid, gas, vapor

Language Objectives:

• Explain the regions of a phase diagram and the relationship between each region and the temperature and pressure of the substance..

Notes:

The phase of a substance (solid, liquid, gas) depends on its temperature and pressure.

phase diagram: a graph showing the phase(s) present at different temperatures and pressures.







Big Ideas	Details	C C	Unit: Matter
		Homework Problems	
	Answer these qu	uestions based on the phase diagrams for wate	er and carbon dioxide.
	1. Approxim temperat	nately what pressure would be necessary to bo ture of 350 K?	il water at a
	2. What is t 350 K?	he minimum pressure necessary for water to e	exist as a liquid at
	3. At approx 10 bar?	ximately what temperature would water boil if	the pressure is
	4. What is t liquid?	he highest temperature at which carbon dioxic	de can exist as a
	5. At 1.0 ba sublimes	ar of pressure, what is the temperature at whic s?	ch carbon dioxide
	6. At room t liquid car	temperature (25 °C ≈ 300 K), what is the minim ′bon dioxide can exist?	um pressure at which
	7. Describe to 400 K a	the phase transitions and temperatures for wa at a pressure of 0.1 bar.	ater going from 200 K
	8. Describe 200 K to 3	the phase transitions and temperatures for ca 300 K at a pressure of 10 bar.	rbon dioxide going

Phases & Phase Changes

Unit: Matter

Details

MA Curriculum Frameworks (2016): HS-PS1-3, HS-PS2-8(MA)

Mastery Objective(s): (Students will be able to...)

• Compare observable states of matter and phase transitions with behavior at the molecular level.

Success Criteria:

- Descriptions include connectedness and motion of molecules.
- Descriptions include comparative descriptions of molecular speed.
- Descriptions relate molecular motion and speed to temperature.

Tier 2 Vocabulary: phase, solid, liquid, gas, vapor

Language Objectives:

• Explain phase changes in terms of changes in molecular behavior.

Notes:

macroscopic: objects or bulk properties of matter that we can observe directly.

microscopic: objects or properties of matter that are too small to observe directly.

In chemistry, the *macroscopic* properties of a substance are determined by *microscopic* interactions between the individual molecules.*

^{*} In this text, the term "molecules" is frequently used to refer to the particles that make up a substance. A molecule is more properly a group of atoms that are covalently bonded together. A substance can be made of individual atoms, molecules, crystals, or other types of particles. This text uses the term "molecules" because the term gives most students a reasonably correct picture of entities that are firmly attached to each other and cannot be pulled apart by physical means.

Big Ideas

States of Matter

The following table shows interactions between the molecules and some observable properties for solids, liquids and gases. Note that the table includes heating curves, which will be discussed in more detail later in the course. For now, understand that a heating curve shows how the temperature changes as heat is added. Notice in particular that the temperature stays constant during melting and boiling.



Phases & Phase Changes

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Unit: Matter

Details

MA Curriculum Frameworks (2016): HS-PS1-3, HS-PS2-8(MA)

Mastery Objective(s): (Students will be able to...)

- Explain conservation of mass, energy, and charge.
- Apply conservation of mass, energy, and charge to situations.

Success Criteria:

- Explanations account for mass, energy, and charge in a variety of changing situations.
- Students can set up and solve equations of the form *initial* + *change* = *final* in the context of problems involving conservation of mass, energy, and charge.

Tier 2 Vocabulary: conservation

Language Objectives:

• Explain what happens to mass, energy and charge in situations where these quantities are transferred from one object or system to another.

Notes:

<u>conservation law</u>: a statement that says that a quantity is the same before and after some change; a "before = after" statement.

- <u>Conservation of Mass/Matter</u>: matter (mass) cannot be created or destroyed (except in a nuclear reaction), only changed in form. The *total* mass before any chemical or physical change equals the total mass after the change. Also the mass of *each element* (each kind of atom) before any chemical or physical change equals the mass of that same element afterwards.
- <u>Conservation of Energy</u>: energy cannot be created or destroyed (except in nuclear reactions), only changed in form. The energy that was present before any change equals the energy that is present after the change.
- <u>Conservation of Electrical Charge</u>: electrical charges cannot be created or destroyed, only transferred from one atom/molecule/etc. to another.

Conservation of Mass, Energy and Charge

Big Ideas	Details	, 0,	0	Unit: Matter
	Sample problem:			
	Q: 12.5 g of sodium hydrogen o dilute hydrochloric acid. Th hydroxide, water, and sodi equation:	carbonate is added to a he reaction produces caum chloride, according	beaker containing Irbon dioxide gas to the following c	g 100. g of , sodium :hemical
	NaHC	$O_3 + HCI \rightarrow NaCI + H_2O$	+ CO ₂	
	If the contents of the beak complete, how much CO ₂ g	er have a mass of 108.2 as escaped?	g after the reacti	on is
	A: The total mass of the chemic afterwards. The initial ma also be 112.5 g. If 108.2 g i 112.5 – 108.2 = 4.3 g, must	cals before the reaction ss is $100 + 12.5 = 112.5$ is still in the beaker, the be the CO ₂ that escape	must equal the to g. The mass after in the remaining r ed.	otal mass rwards must mass,
	Nuclear Reactions			
	In a nuclear reaction, mass can	be converted to energy	/ according to the	e formula:
		$E = mc^2$		
	Where:			
	E = energy	<i>m</i> = mass	<i>c</i> = the sp	eed of light
	Note that the speed of light is a even larger: $9.00 \times 10^{16} \frac{m^2}{s^2}$. Th extremely large amount of ene comes from—mass that was tu	a very large number: 3 is means that a very sm rgy. This is where the e rned into energy.	$00 \times 10^8 \frac{m}{s}$. There all amount of mathered energy in a nuclea	efore, c ² is ss creates an r explosion
	This law is called the law of <u>Cor</u> equation, $E = mc^2$) describes th the only way we know of that e	nservation of Mass and e only we know of that energy can be created.	<u>Energy</u> . This law mass can be dest	(and its royed, and

Conservation of Mass, Energy and Charge Page: 120

Big Ideas	Details Unit: Matter
	Homework Problems
	 Suppose your breakfast contained 500 Calories of energy. Suppose you missed your bus (or your ride) and you had to walk 2 miles to school, which burned 200 Calories. How many Calories of energy from your breakfast are left for you to get through your morning classes?
	Answer: 300 Calories
	2. In your car's gas tank, the following chemical reaction occurs:
	$C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$
	The gasoline (C_8H_{18}) in a typical car's fuel tank weighs about 80 pounds. Burning that much gas uses about 300 pounds of oxygen from the Earth's atmosphere, and it produces 120 pounds of H ₂ O. How many pounds of CO ₂ did that tank of gas produce?
	Answer: 260 lbs. 3. A nuclear reactor converts 4 × 10 ⁻⁹ kg of uranium to energy. How much energy is produced?
	Answer: 3.6×10^8 J or 360 MJ

Details Un	it: Gases
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Gas Conversion Factors	125
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Ideal Gas Law	137
Partial Pressures	141
Molecular Speed & Effusion	148
Standards addressed in this chapter:	
Massachusetts Curriculum Frameworks & Science Practices (2016):	
Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS2-8(MA) Use kinetic molecular theory to compare the streng electrostatic forces and the prevalence of interactions that occ molecules in solids, liquids, and gases. Use the combined gas la determine changes in pressure, volume, and temperature in gr	
	Details Unit Introduction: Gases Unit: Gases Topics covered in this chapter: Gases & Kinetic-Molecular Theory Gas Conversion Factors Gas Laws Ideal Gas Law Partial Pressures Molecular Speed & Effusion Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS2-8(MA) Use kinetic molecular theory to compare the strengths o electrostatic forces and the prevalence of interactions that occur be molecules in solids, liquids, and gases. Use the combined gas law to determine changes in pressure, volume, and temperature in gases.

Unit: Gases

Details

MA Curriculum Frameworks (2016): HS-PS2-8(MA)

Mastery Objective(s): (Students will be able to...)

• Explain how each aspect of Kinetic-Molecular Theory applies to gases.

Success Criteria:

- Descriptions account for behavior at the molecular level.
- Descriptions account for measurable properties, *e.g.*, temperature, pressure, volume, *etc.*

Tier 2 Vocabulary: kinetic, gas, ideal, real

Language Objectives:

• Explain how gas molecules behave and how their behavior relates to properties we can measure.

Notes:

Recall the following definitions:

- solid: molecules rigidly bonded (definite shape & volume)
- <u>liquid</u>: molecules bonded (definite volume), but loosely. Bonds continually breaking & re-forming (indefinite shape)
- gas: molecules not bonded (indefinite volume & shape)
- <u>plasma</u>: heat of surroundings > ionization energy, so electrons are loosely bonded & continually dissociate from and re-associate with ions. Electrical charge is fluid and in continual motion.

evaporation: conversion of liquid to gas.

boiling point: the temperature at which a liquid completely evaporates.

<u>normal boiling point</u>: the boiling point of a liquid when the pressure is 1 atm (average atmospheric pressure at sea level).

Details Unit: Gases **Big Ideas Kinetic-Molecular Theory** Kinetic-Molecular Theory (KMT) is a theory, developed by James C. Maxwell and Ludwig Boltzmann, that predicts the behavior of gases by modeling them as moving molecules. The theory states that: · Gases are made of very large numbers of molecules Molecules are constantly moving (obeying Newton's laws of motion), and their speeds are constant • Molecules are very far apart compared with their diameter • Molecules collide with each other and walls of container in elastic collisions • Molecules behaving according to KMT are not reacting^{*} or exerting any other forces (attractive or repulsive) on each other. temperature: a measure of the average kinetic energy of the molecules of a substance ideal gas: a gas whose molecules behave according to KMT. Most gases are ideal under some conditions (but not all). real gas: a gas that is not behaving according to KMT. This can occur with all gases, most commonly at temperatures and pressures that are close to the solid or liquid sections of the phase diagram for the substance. Liquid Solid Pressure Gas Temperature Of course, reactions can occur, but chemical reactions are part of collision theory, which is separate from KMT.

Gases & Kinetic-Molecular Theory

Big Ideas

Details

Measurable Properties of Gases

All gases have the following properties that can be measured:

Property	Variable	Units	Description
amount	n	moles (mol)	amount of gas (1 mol = 6.02 × 10 ²³ molecules)
volume	V	liters (L)	space that the gas takes up
temperature	Т	kelvin (K)	ability to transfer heat (proportional to the average kinetic energy of the molecules)
pressure	pressure P		average force on the walls of the container due to collisions between the molecules and the walls

Notes about calculations:

- Temperature must be absolute, which means you <u>must</u> use Kelvin. A temperature of 0 in a gas laws calculation can only mean absolute zero.
- Pressures must be absolute. (For example, you can't use a tire gauge because it measures "gauge pressure," which is the difference between atmospheric pressure and the pressure inside the tire.) A pressure of 0 in a gas laws calculation can only mean that there are no molecules colliding with the walls.

Gas Conversion Factors

Unit: Gases

MA Curriculum Frameworks (2016): HS-PS2-8(MA)

Mastery Objective(s): (Students will be able to...)

• Choose conversion factors based on the units needed for a calculation or conversion.

Success Criteria:

• Conversion factor has the same units as other numbers in a chosen word problem or situation.

Tier 2 Vocabulary: conversion, absolute, standard, vacuum

Language Objectives:

• Explain and defend the choice of a conversion factor or constant for use in a problem involving gases.

Notes:

<u>absolute zero</u>: the temperature at which molecules are moving so slowly that they can't transfer energy to other molecules. Absolute zero is -273.15 °C = 0 K

vacuum: the absence of gas molecules. In a total vacuum, the Pressure = 0

```
"Standard Pressure" = 1 bar*
```

"Standard Temperature" = 0°C = 273.15 K

S.T.P. ("Standard Temperature and Pressure") = 0 °C and 1 bar.

"Room Temperature" = 25 °C = 298 K

1 mole of an ideal gas has a volume of 22.7 L at S.T.P.

^{*} In 1982, the IUPAC defined standard pressure to be exactly 1 bar (= 100 kPa = 0.987 atm). However, many chemists and many standardized assessments still use 1 atm.

Details					Unit: Gas
		Conver	sion Fa	ctors	
Pressure:					
1 atm	= 101.325 = 101.325	$kPa \equiv 0.10132$ $\frac{N}{m^2} \equiv 101325$	25 MPa≡1 Pa	.01325 bar	
1 atm	≡ 760 mm	Hg≡760 torr	= 29.92 in.	Hg	
1 atm	= 14.696 <u> </u> ir	<u>b.</u> . ² = 14.696 p	si ("psi" = "	pounds per squ	are inch")
Volume:					
$1 \text{ mL} \equiv 1 \text{ cm}^3$					
$1 L \equiv 1000 m^2$	3				
Moles:					
1 mol = 22.7 L at S.T.P.*					
Use dimensional analysis to turn the molar mass of a compound (measured in $\frac{g}{m}$				(measured in $\frac{g}{mol}$)	
into a conversion factor between grams and moles.					
Temperature:					
$K \equiv {}^{\circ}C + 273.2$	15	°F≡(1.8×°C) + 32	$^{\circ}R \equiv ^{\circ}F + ^{\circ}$	459.67
The Gas Co	onstant:				
The gas constant R is a natural constant that appears in several relationships in chemistry, including the ideal gas law (which we will study in a subsequent class).					elationships in Ibsequent class).
$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$					
$R = 8.31 \frac{L \cdot kPa}{mol \cdot K} = 8.31 \frac{J}{mol \cdot K} = 8.31 \times 10^{-3} \frac{kJ}{mol \cdot K}$					
R = 6	52.4 L·torr mol·K		R=1.987	$\frac{\text{cal}}{\text{mol}\cdot\text{K}} \equiv 1.987 \frac{1}{\text{lb}\cdot\text{K}}$	BTU -mol· [°] R
1					

 * Massachusetts assessments still use the outdated definition of S.T.P. The volume of one mole of an ideal gas at 1 atm and 0 $^{\circ}$ C is 22.4 L.

Use this space for summary and/or additional notes:

Big Ideas

Gas Laws

Unit: Gases

Details

Big Ideas

MA Curriculum Frameworks (2016): HS-PS2-8(MA)

Mastery Objective(s): (Students will be able to...)

- Qualitatively describe the relationship between any two of the quantities: *number of particles, temperature, pressure,* and *volume* in terms of Kinetic Molecular Theory (KMT).
- Quantitatively determine the *number of particles, temperature, pressure,* or *volume* in a before & after problem in which one or more of these quantities is changing.

Success Criteria:

- Descriptions relate behavior at the molecular level to behavior at the macroscopic level.
- Solutions have the correct quantities substituted for the correct variables.
- Chosen value of the gas constant has the same units as the other quantities in the problem.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: ideal, law

Language Objectives:

- Identify each quantity based on its units and assign the correct variable to it.
- Understand and correctly use the terms "pressure," "volume," and "temperature," and "ideal gas."
- Explain the placement of each quantity in the ideal gas law.

Labs, Activities & Demonstrations:

- Vacuum pump (pressure & volume) with:
 - \circ balloon (air vs. water)
- \circ shaving cream
- Absolute zero apparatus (pressure & temperature)
- Can crush (pressure, volume & temperature)

Big Ideas	Details	Unit: Gases
	Notes:	
	ideal gas: a gas that behaves as if each molecule acts indepen- kinetic-molecular theory. Specifically, this means the mol and move freely in straight lines at constant speeds. Whe collide, the collisions are perfectly elastic, which means th other with no energy or momentum lost.	dently, according to ecules are far apart, in the molecules iey bounce off each
	Most gases behave ideally except at temperatures and provaporization curve on a phase diagram. (<i>I.e.,</i> gases stop b conditions are close to those that would cause the gas to solid.)	essures near the ehaving ideally when condense to a liquid or
	Ogo Solid Liquid Juinteen and state Gas Temperature	

Big Ideas	Details Unit: Gases
	Note about Subscripts and Variables
	When variables appear more than once in an equation with different values each time, we use subscripts to group them. You have already seen this a few times in math, <i>e.g.</i> , in the formula for the slope of a line and the distance formula:
	slope: $m = \frac{y_2 - y_1}{x_2 - x_1}$ distance: $d = \sqrt{(x_2 - x_1)^2 + (y_2 - y_1)^2}$
	In the above examples, the subscripts "1" and "2" are used to group the x and y values based on whether they refer to the first point (x_1, y_1) , or the second one (x_2, y_2) .
	In chemistry, we use subscripts the same way. For example, if a gas is heated, that means the temperature is changing. We refer to the starting temperature (temperature #1) as T_1 , and the ending temperature (temperature #2) as T_2 . The same concept applies to other variables as well, such as moles (<i>n</i>), volume (<i>V</i>), and pressure (<i>P</i>).
	Proportionalities
	directly proportional: if two quantities are <u>directly</u> proportional, as one increases, the other increases proportionately.
	If x and y are directly proportional, then $x \propto y$ which means $x = ky$ and $\frac{x}{y} = k$
	where <i>k</i> is a constant. You should notice that <i>x</i> and <i>y</i> are either <u>numerator and</u> <u>denominator</u> in a fraction, or are on <u>opposite sides</u> of the equals sign.
	inversely proportional: if two quantities are inversely proportional, as one increases, the other decreases proportionately.
	If x and y are inversely proportional, then $x \propto \frac{1}{y}$ which means $xy = k$ where k is
	a constant. You should notice that <i>x</i> and <i>y</i> are on the <u>same side</u> of the equals sign.

Gas Laws

Use this space for summary and/or additional notes:

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Details

Big Ideas

Avogadro's Principle

In 1811, Italian physicist Amedeo Avogadro (whose full name was Lorenzo Romano Amedeo Carlo Avogadro di Quaregna e di Cerreto) published the principle that equal volumes of an ideal gas at the same temperature and pressure must contain equal numbers of particles.

What did we do?	What happened?	What are the molecules doing?	Conclusion
put more (moles of) air into a balloon $n \uparrow$	the volume of the balloon got larger V↑	crowding each other → pushing each other farther away	<i>n</i> and <i>V</i> are directly proportional. $\frac{V}{n}$ = constant

If the pressure and temperature are constant, then for an ideal gas:

$$\frac{V_1}{N_1} = \frac{V_2}{N_2}^*$$

Because it is almost always more convenient to work with moles of gas (n) rather than particles (N), we can rewrite Avogadro's principle as:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

* Avogadro's principle is usually stated $\frac{n_1}{V_1} = \frac{n_2}{V_2}$. I have inverted it in these notes so that the quantities in the numerator and denominator are the same as the quantities in the

numerator and denominator of the combined gas law.

Big Ideas

Boyle's Law

In 1662, British physicist and chemist Robert Boyle published his findings that the pressure and volume of a gas were inversely proportional.

What did we do?	What happened?	What are the molecules doing?	Conclusion
decrease pressure by putting a balloon in a vacuum chamber P↓	the volume of the air inside the balloon increased $v \uparrow$	colliding with less force → pushing each other less far away	P and V are inversely proportional. PV = constant

Therefore, if the temperature and the number of particles of gas are constant, then for an ideal gas:

 $P_1V_1 = P_2V_2$

(Note that by convention, gas laws use subscripts "1" and "2" instead of "i" and "f".)

Charles' Law

In the 1780s, French physicist Jacques Charles discovered that the volume and temperature of a gas were directly proportional.

What did we do?	What happened?	What are the molecules doing?	Conclusion
cool gas by putting a soda can full of very hot air into cool water $T \downarrow$	the volume of the air got smaller and crushed the can $V\downarrow$	moving more slowly → pushing each other less far away	V and T are directly proportional. $\frac{V}{T} = \text{constant}$

If pressure and the number of particles are constant, then for an ideal gas:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Gay-Lussac's Law

In 1702, French physicist Guillaume Amontons discovered that there is a relationship between the pressure and temperature of a gas. However, precise thermometers were not invented until after Amontons' discovery, so it wasn't until 1808, over a century later, that French chemist Joseph Louis Gay-Lussac confirmed this law mathematically. The pressure law is most often attributed to Gay-Lussac, though some texts refer to it as Amontons' Law.

What did we do?	What happened?	What are the molecules doing?		
increase temperature by heating a metal sphere full of air $T\uparrow$	the pressure of the air increased ₽↑	moving faster → colliding with more force	<i>P</i> and <i>T</i> are directly proportional. $\frac{P}{T} = \text{constant}$	

If volume and the number of particles are constant, then for an ideal gas:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

The Combined Gas Law

We can combine each of the above principles. When we do this (keeping P and V in the numerator and n and T in the denominator for consistency), we get following relationship for an ideal gas:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} = \text{constant}$$

Note, however, that in most problems, the number of moles of gas remains constant. This means $n_1 = n_2$ and we can cancel it from the equation, which gives:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This equation is called the "combined gas law", which is used to solve most "before/after" problems involving ideal gases.

Use this space for summary and/or additional notes:

Big Ideas

Details

Big Ideas	Details Unit: Gases
	When using the combined gas law, any quantity that is not changing may be cancelled out of the equation. (If a quantity is not mentioned in the problem, you can assume that it is constant and may be cancelled.)
	This brings us to an important point about science problems: <u>If something is not</u> <u>mentioned in a problem</u> , always assume that it doesn't affect the problem. On a standardized test like MCAS or an AP [®] test, it's usually best to state those assumptions explicitly, because if your assumption is valid and you do the rest of the problem correctly, you will almost always receive some credit, even if your assumption was different from what the person who wrote the problem intended.
	For example, suppose a problem doesn't mention anything about temperature. That means T is constant and you can cancel it. When you cancel T from both sides of the combined gas law, you get:
	$\frac{P_1V_1}{\mathcal{X}_1} = \frac{P_2V_2}{\mathcal{X}_2}$ which simplifies to $P_1V_1 = P_2V_2$ (Boyle's Law)
	Solving Problems Using the Combined Gas Law
	You can use this method to solve any "before/after" gas law problem:
	1. Determine which variables you have
	2. Determine which values are <i>initia1</i> (#1) vs. <i>final</i> (#2).
	 Start with the combined gas law and cancel any variables that are explicitly not changing or omitted (assumed not to be changing).
	 Substitute your numbers into the resulting equation and solve. (Make sure all initial and final quantities have the same units, and don't forget that temperatures <u>must</u> be in Kelvin!)

Big Ideas	Details			Unit: Gase		
	Samp	le problem:				
	Q: A gas has a temperature of 25 °C and a pressure of 1.5 bar. If the gas is heater to 35 °C, what will the new pressure be?					
	A: 1.	Find which variables we have.				
		We have two temperatures (25 °C and and the new pressure that we're looki	35 °C), and two pressures ng for).	(1.5 bar		
	2.	Find the action being done on the gas about the gas <i>before</i> the action is time the gas <i>after</i> the action is time "2".	("heated"). Anything that "1", and anything that is t	was true rue about		
		<u>Time 1 ("before")</u> :	<u>Time 2 ("after")</u> :	:		
		<i>P</i> ₁ = 1.5 bar	$P_2 = P_2$			
		<i>T</i> ₁ = 25 °C + 273 = 298 K	<i>T</i> ₂ = 35 °C + 273	= 308 K		
	3.	Set up the formula. We can cancel vol doesn't mention it:	ume (<i>V</i>), because the prob	lem		
		$\frac{P_1 Y_1}{T_1} = \frac{P_2 Y_2}{T_2}$ which gives us $\frac{P_1}{T_1} =$	$=\frac{P_2}{T_2}$ (Gay-Lussac's Law)			
	4.	Plug in our values and solve: $\frac{1.5 \text{ bar}}{298 \text{ K}} = \frac{P_2}{308 \text{ K}}$	\rightarrow $P_2 = 1.55$ bar			

Use this space for summary and/or additional notes:

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Gas Laws

Big Ideas	Details Unit: Gases
	Homework Problems
	Solve these problems using one of the gas laws in this section. Remember to convert temperatures to Kelvin!
	 A sample of oxygen gas occupies a volume of 250. mL at a pressure of 740. torr. What volume will it occupy at 800. torr?
	Answer: 231.25 mL
	 A sample of O₂ is at a temperature of 40.0 °C and occupies a volume of 2.30 L. To what temperature should it be raised to occupy a volume of 6.50 L?
	Answer: 612 °C
	3. H ₂ gas was cooled from 150. °C to 50. °C. Its new pressure is 750 torr. What was its original pressure?
	Answer: 980 torr
	4. A 2.00 L container of N₂ had a pressure of 3.20 atm. What volume would be necessary to decrease the pressure to 98.0 kPa?
	(Hint: notice that the pressures are in different units. You will need to convert one of them so that both pressures are in either atm or kPa.)
	Answer: 6.621

	Gas Laws	Page: 136	
Big Ideas	Details	Unit: Gases	
	5. A sample of air has a volume of 60.0 mL at S.T.P. What volume sample have at 55.0 °C and 745 torr?	will the	
	Answer: 73.5 mL		
	6. N₂ gas is enclosed in a tightly stoppered 500. mL flask at 20.0 °C The flask, which is rated for a maximum pressure of 3.00 atm, i 680. °C. Will the flask explode?	C and 1 atm. s heated to	
	Answer: $P_2 = 3.25$ atm. Yes, the flask will explode.		
	7. A scuba diver's 10. L air tank is filled to a pressure of 210 bar at temperature of 32.0 °C. When the diver is breathing the air un water temperature is 8.0 °C, and the pressure is 2.1 bar.	a dockside: derwater, the	
	a. What volume of air does the diver use?		
	Answer: 921 L		
	b. If the diver uses air at the rate of 8.0 L/min, how long will last?	the diver's air	
	Answer: 115 min		

Ideal Gas Law

Unit: Gases

MA Curriculum Frameworks (2016): HS-PS2-8(MA)

Mastery Objective(s): (Students will be able to...)

- Describe the relationship between any two variables in the ideal gas law.
- Apply the ideal gas law to problems involving a sample of gas.

Success Criteria:

- Solutions have the correct quantities substituted for the correct variables.
- Chosen value of the gas constant has the same units as the other quantities in the problem.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: ideal, law

Language Objectives:

- Identify each quantity based on its units and assign the correct variable to it.
- Explain the placement of each quantity in the ideal gas law.

Notes:

ideal gas: a gas that behaves according to Kinetic-Molecular Theory (KMT).

When we developed the combined gas law, before we cancelled the number of moles, we had the equation:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} = \text{constant}$$

Because *P*, *V*, *n* and *T* are all of the quantities needed to specify the conditions of an ideal gas, this expression must be true for *any ideal gas* under *any conditions*. If *V* is in liters, *P* is in kPa, *n* is in moles and *T* is in Kelvin, then the value of this constant is $8.31 \frac{\text{L-kPa}}{\text{mol-K}}$. This number is called "the gas constant", and is denoted by the variable *R* in equations.

Therefore, we can rewrite the above expression as:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} = R \text{ and therefore } P_1V_1 = n_1RT_1 \text{ and } P_2V_2 = n_2RT_2$$

Because this expression is true under any conditions, 1, 2, or whatever, we can drop the subscripts:

PV = nRT

Big Ideas	Details Unit: Ga	ises		
	Choosing a value of R			
	The purpose of the gas constant R is to convert the quantity nT from units of mol to units of pressure × volume. As we saw earlier in this chapter, the gas constant has different values, depending on the units it needs to cancel:	٠K		
	$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$			
	$R = 8.31 \frac{L \cdot kPa}{mol \cdot K} = 8.31 \frac{J}{mol \cdot K} = 8.31 \times 10^{-3} \frac{kJ}{mol \cdot K}$			
	$R = 62.4 \frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}} \qquad R = 1.987 \frac{\text{cal}}{\text{mol} \cdot \text{K}} = 1.987 \frac{\text{BTU}}{\text{lb} - \text{mol}^{\circ} \text{R}}$			
	In order for the units in a problem to work out properly, <u>R needs to have exactly to same units as the values in the problem</u> . In the problems we will be solving, this usually means you need to look at the pressure units and <u>choose the value of R the has the same pressure unit as the pressure given in the problem</u> .	<u>he</u> hat		
	Amazing Fact:			
	In the metric system, the unit of pressure (the Pascal) is a combination of the S.I. units $\frac{\text{kg}}{\text{m} \cdot \text{s}^2}$, and volume has units of m ³ . This means that pressure times volume			
	(<i>PV</i>) has S.I. units of $\frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}$, which happens to equal Joules (the unit for energy).			
	Dimensional analysis tells us that if the units work, the resulting formula must be correct.			
	So, if <i>PV</i> is equivalent to energy, <i>nRT</i> <u>must also be</u> equivalent to energy. In fact, <i>I</i> is the energy that the gas expends in doing work on its surroundings (such as by expanding a balloon or pushing on a piston), and <i>nRT</i> is the kinetic energy of the individual gas molecules.	י∕		
	In other words, the ideal gas law <i>PV</i> = <i>nRT</i> is simply the law of conservation of energy, applied to gases!			

	lueal Gas L	-dW	Page: 139			
Big Ideas	Details		Unit: Gases			
	Solving Problems Using the Ideal Gas Law If a gas behaves according to the ideal gas law, simply substitute the values for pressure, volume, number of moles (or particles), and temperature into the equation. Be sure your units are correct (especially that temperature is in Kelvin), and that you use the correct constant, depending on whether you know the number of particles or the number of moles of the gas.					
	Sample Problem:					
	n	Р				
	A 3.50 mol sample of an ideal gas has a 35 °C. What is its volume? $T \rightarrow K$ V	sample of an ideal gas has a pressure of 1.20 atm and a temperature of at is its volume? v				
	Answer:					
	First, we need to declare our variab units:	les and choose a value for R base	ed on the			
	<i>P</i> = 1.20 atm	<i>n</i> = 3.50 mol				
	V = V (because we don't' know it	$R = 0.0821 \frac{\text{L-atm}}{100}$ (Choose	this one			
	yet)	because <i>P</i> is in atm.)				
		<i>T</i> = 35 °C + 273 = 308 К				
	Then we substitute these numbers into	the ideal gas law and solve:				
	PV	√ = n RT				
	(1.20 atm) V = (3.50	nol)(0.0821 <u>L-atm</u>)(308 K)				
	1.20	0 <i>V</i> = 88.5				
	88	88.5				
	$V = \frac{1}{1.2}$	=/3./5L 20				
	Use this space for summary and/or add	itional notes:				

Ideal Gas Law

Big Ideas	Details	Unit: Gase
	Homework Problems	
	Use the ideal gas law to solve the following problems. Be sure to choose appropriate value for the gas constant and to convert temperatures to Ke	the lvin.
	 A sample of 1.00 moles of oxygen at 50.0 °C and 98.6 kPa occupies volume? 	what
	Answer: 27.2 L	
	2. If a steel cylinder with a volume of 1.50 L contains 10.0 moles of os under what pressure is the oxygen if the temperature is 27.0 °C?	kygen,
	Answer: 164 atm = 125,000 torr = 16,600 kPa	
	3. In a gas thermometer, the pressure of 0.500 L of helium is 113.30 k	<pa a<="" at="" td=""></pa>
	temperature of −137 °C. How many moles of gas are in the samp	le?
	Answer: 0.050 mol	
	4. A sample of 4.25 mol of hydrogen at 20.0 °C occupies a volume of 2 Under what pressure is this sample?	25.0 L.
	Answer: 4.09 atm = 3 108 torr = 414 kPa	

Partial Pressures

Unit: Gases

Details

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to ...)

• Calculate partial pressures based on conservation of matter.

Success Criteria:

- Solutions have the correct quantities substituted for the correct variables.
- Mole fractions are paired correctly with their partial pressures.
- If the problem requires the ideal gas law, chosen value of the gas constant has the same units as the other quantities in the problem.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: mole

Language Objectives:

• Describe the pairing of each gas with its mole fraction and pressure.

Notes:

<u>Partial Pressure</u>: the partial pressure of a gas is the amount of pressure that would result from *only* the molecules of that gas. The partial pressure of a substance is denoted by the variable *P* (for pressure) and the chemical formula of the substance as a subscript. For example, the partial pressure of carbon dioxide in a sample would be denoted by P_{CO_3} .

<u>Dalton's Law of Partial Pressures</u>: the sum of all of the partial pressures in a sealed container equals the total pressure.

$$P = P_T = P_1 + P_2 + P_3 + \dots$$

(To make things more clear, we will use P_T to mean the total pressure.)

<u>mole fraction</u> (χ): the fraction of the total moles (or molecules) that are the compound of interest. For example, if we have 20 moles of gas, and 9 moles are N₂, the mole fraction of N₂ is:

$$\chi_{N_2} = \frac{9 \,\text{mol}\,N_2}{20 \,\text{mol}\,\text{total}} = 0.45$$

Big Ideas

Details

Partial Pressures

Suppose we had the following tank, with a total pressure of 1.00 atm:

$$\begin{bmatrix} & & & & & O_2 & & \\ & N_2 & O_2 & N_2 & & N_2 & \\ & O_2 & & & & & O_2 & \\ & O_2 & & N_2 & & O_2 & \\ & N_2 & O_2 & & N_2 & \\ & N_2 & O_2 & & O_2 & O_2 \end{bmatrix}$$

If we ignore all of the molecules except for nitrogen, the tank would look like this:

 N_2

 N_2

 N_2

 N_2

 N_2



 $P_{N_2} = \chi_{N_2} P_T$

$$P_{N_2} = (0.45)(1 \text{ atm}) = 0.45 \text{ atm}$$

Similarly, because 55 % of the molecules are oxygen, this means:

 N_2

 N_2

 $P_{O_2} = \chi_{O_2} P_T$ $P_{O_2} = (0.55)(1 \text{ atm}) = 0.55 \text{ atm}$

Note that the two partial pressures add up to the total pressure:

$$P_{\rm T} = P_{\rm N_2} + P_{\rm O_2} = 0.45 \text{ atm} + 0.55 \text{ atm} = 1 \text{ atm}$$

$$\chi_{0_2} = 0.55$$

Partial Pressures

Using Dalton's Law with the Ideal Gas Law

Recall the two tanks from our example. Assuming N_2 and O_2 are behaving like ideal gases, the ideal gas law must be true in both tanks.



In other words, the ideal gas law can be used either with the total moles and total pressure, or with the moles of one specific gas and the partial pressure of that gas.

Use this space for summary and/or additional notes:

Big Ideas

Details

Big Ideas	Details							Unit:
	Vapor Pressure							
	vapor pressure (P_{vap}) the partial pressure of a substance due to evaporate					oration.		
	Because liquids are continually forming and breaking bonds, when a liquid n at the surface breaks its bonds with other liquids, it can escape the attractiv of the other liquid molecules and become a vapor molecule. The tendency molecules to do this, when expressed as a partial pressure, is called the vapor pressure.						iquid mo tractive f dency for he vapor	
	Vapor pressure is a function of the kinetic energy of the molecules, which me vapor pressure increases with temperature. At the boiling point, all of the molecules have enough energy to enter the gas phase. This means that at the boiling point, the vapor pressure must be equal to the ambient (atmospheric) pressure.					hich mea of the nat at the ospheric)		
	The following ta	able sho	ws the vap	or press	ure of wa	ter at di	fferent te	mperatur
			Vapo	or Press	ure of Wa	iter		
		Temp (°C)	P _{vap} (kPa)	Temp (°C)	P _{vap} (kPa)	Temp (°C)	P _{vap} (kPa)	
		0.01	0.61173	30	4.2455	70	31.176	
		1	0.65716	35	5.6267	75	38.563	
		4	0.81359	40	7.3814	80	47.373	
		5	0.87260	45	9.5898	85	57.815	
		10	1.2281	50	12.344	90	70.117	
		15	1.7056	55	15.752	95	84.529	
		20	2.3388	60	19.932	100	101.32	
		25	3.1691	65	25.022	105	120.79	
	Relative humidi vapor pressure.	ty is the	actual par	tial pres	sure of w	ater in a	ir as a per	centage
	For example, su pressure of air a would be 50 %.	ippose a at 30 °C i	ir at 30 °C is 5.6 kPa.	(86 °F) ha 2.8 kPa	as a partia is half of !	al pressu 5.6 kPa, s	ire of 2.8 so the rel	kPa. The ative hun

lecule forces

of its

vapor nidity
		i ai tiai i i Cosuit	- 3 Fage. 14		
Big Ideas	Details		Unit: Gase		
	Sample problem	า:			
	A 12.0 L tank of gas the partial pressure are in the tank? He	A 12.0 L tank of gas has a temperature of 30.0 °C and a total pressure of 1.75 atm. If the partial pressure of oxygen in the tank is 0.350 atm, how many moles of oxygen are in the tank? How many total moles of gas are in the tank?			
	Solution:				
	For oxygen:				
		<i>P</i> ₀₂ = 0.350 atm	n = n		
		V = 12.0 L	$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$		
			T = 30.0 °C + 273 = 303.0 K		
		$P_{O_2}V = n$	_{o2} R <i>T</i>		
		(0.350)(12.0) = <i>n</i> ₀₂	(0.0821)(303.0)		
		n ₀₂ = 0.16	59 mol		
	You could figure ou the total moles:	it the total moles two wa	ys. One is to use the ideal gas law on		
		<i>P</i> = 1.75 atm	n = n		
		V = 12.0 L	$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$		
			T = 30.0 °C + 273 = 303.0 K		
		<i>PV</i> = <i>n</i>	R <i>T</i>		
		(1.75)(12.0) = <i>n</i> (0).0821)(303.0)		
		<i>n</i> = 0.844	1 mol		

Partial Pressures

Big Ideas	Details Unit: Gases
	The other way to find the total moles is to use the mole fraction and the partial pressure:
	$P_{O_2} = \chi_{O_2} P_T$
	We know that
	<i>P</i> ₀₂ = 0.350 atm
	<i>P</i> _T = 1.75 atm
	0.350 atm = $\chi_{0_2}(1.75 \text{ atm})$
	$\chi_{\rm O_2} = \frac{0.350{\rm atm}}{1.75{\rm atm}} = 0.200$
	Now that we know the mole fraction of O_2 , we can figure out the total moles:
	$\chi_{O_2} = \frac{n_{O_2}}{n_{T}}$ $0.200 = \frac{0.169 \text{ mol } O_2}{n_{T}}$ $n_{T} = \frac{0.169}{0.200} = 0.845 \text{ mol}$
	Homework Problems
	 A 5 L container contains 0.125 mol of O₂ and 1.000 mol of He at 65 °C. What is the partial pressure of each gas? What is the total pressure?
	Answer: 6.24 atm

Partial Pressures

Big Ideas	Details Unit: Gases
	2. A 50 L gas cylinder contains 186 mol of N ₂ and 140 mol of O ₂ . If the temperature is 24 °C, what is the total pressure in the cylinder?
	Answer: 159 atm 3. A sample of O ₂ gas is collected by water displacement at 25 °C. If the atmospheric pressure in the laboratory is 100.7 kPa and the vapor pressure of water is 3.17 kPa at 25 °C, what is the partial pressure of the O ₂ gas in the sample?
	 Answer: 97.5 kPa 4. Two flasks are connected with a stopcock. The first flask has a volume of 5 liters and contains nitrogen gas at a pressure of 0.75 atm. The second flask has a volume of 8 ℓ and contains oxygen gas at a pressure of 1.25 atm. When the stopcock between the flasks is opened and the gases are free to mix, what will the (total) pressure be in the resulting mixture?
	Answer: 1.058 atm

Molecular Speed & Effusion

Unit: Gases

Details

MA Curriculum Frameworks (2016): N/A (optional topic)

Mastery Objective(s): (Students will be able to...)

• Determine the velocities of gas molecules from their molar masses.

Success Criteria:

- Solutions use the equation appropriate for the information given.
- Solutions have the correct quantities substituted for the correct variables.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: root, mean, square

Language Objectives:

• Describe the pairing of each gas with its mole fraction and pressure.

Notes:

effusion: the escape of a gas through a small opening.

- <u>Graham's Law of Effusion</u>: lighter molecules move faster and heavier molecules move more slowly. If you have a small opening, such as a leak in a tank, the lighter molecules will escape faster.
- <u>temperature</u>: a measure of the average kinetic energy of the molecules in a substance
- <u>root mean square speed</u>: the average speed of the gas molecules. "Root mean square" means this average is calculated by determining the average of the squares of the speeds, then taking the square root of the result.

Root Mean Square Speed of Gas Molecules

The root mean square velocity (v_{rms}) of gas molecules is given by the formula^{*}:

$$v_{rms} = \sqrt{\frac{3000\,RT}{M}}$$

where *M* is the molar mass of the gas in $\frac{g}{mol}$, *R* is the gas constant (8.31 $\frac{J}{mol\cdot K}$), and *T* is the temperature in Kelvin. The quantity v_{rms} comes out in units of $\frac{m}{s}$.

Sample Problem:

Details

Big Ideas

A tank is filled with a mixture of helium and oxygen at a temperature of 25 °C. Calculate the root mean square speed (v_{rms}) of the oxygen molecules. (Note: oxygen has a molar mass of $32 \frac{g}{mol}$.)

$$T = 25^{\circ}\text{C} + 273 = 298 \text{ K} \text{ and } M_{O_2} = 32 \frac{\text{g}}{\text{mol}}$$

$$v_{rms} = \sqrt{\frac{3000 \text{ RT}}{M}}$$

$$v_{rms} = \sqrt{\frac{3000 (8.31 \frac{\text{J}}{\text{mol} \text{K}})(298 \text{ K})}{32 \frac{\text{g}}{\text{mol}}}}$$

$$v_{rms} = \sqrt{232,161} = 482 \frac{\text{m}}{\text{s}}$$

* This formula is usually given as $v_{rms} = \sqrt{\frac{3RT}{M}}$, but this requires that M be expressed in $\frac{\text{kg}}{\text{mol}}$. Because we are not deriving the formula in this class, it makes sense to change the factor from 3 to 3000 in order to keep molar mass expressed in the more familiar units of $\frac{\text{g}}{\text{mol}}$.

Molecular Speed & Effusion

Details	Unit: Gases
Graham's Law	
From physics, the formula for kinetic energy is:	
$E_{kinetic} = \frac{1}{2}mv^2$	
where m is the mass and v is the velocity (speed).	
If two molecules have the same temperature, then they have the same energy. If molecule #1 has mass m_1 and velocity v_1 and molecule #2 h and velocity v_2 , then the kinetic energies are:	e kinetic as mass <i>m</i> ₂
$E_{kinetic} = \frac{1}{2}m_1v_1^2 = \frac{1}{2}m_2v_2^2$	
If we cancel the ½ from both sides, and rearrange so the masses are or the speeds are on the other, we get:	n one side and
$\frac{v_2^2}{v_1^2} = \frac{m_1}{m_2}$	
Taking the square root of both sides:	
$\frac{v_2}{v_1} = \sqrt{\frac{m_1}{m_2}}$	
This formula is called Graham's Law, named after Thomas Graham who proposed it.	o first
For the purpose of these calculations, the mass of a molecule is its molecule is its molecule is simply the same number of grams as the molecule's atomic m	lar mass, ass.
Because the rate of effusion (r) is the velocity divided by time, the unit	s of time
cancel, giving $\frac{r_2}{r_1} = \frac{v_2}{v_1}$	

Use this space for summary and/or additional notes:

Big Ideas

Big Ideas	Details	Unit: Gases
	Sample Problem:	
	If the speed of oxygen molecules in a tank is $482\frac{m}{s}$, calculate the speed of oxygen molecules in a tank is $482\frac{m}{s}$.	peed of helium
	molecules in the same tank. (O ₂ has a molar mass of $32.0 \frac{g}{mol}$ and	helium has a
	molar mass of $4.01 \frac{g}{mol}$.)	
	$M_{O_2} = 32.0 \frac{g}{mol}$ and $M_{He} = 4.01 \frac{g}{mol}$	
	$\frac{v_{\rm He}}{v_{\rm O_2}} = \sqrt{\frac{m_{\rm O_2}}{m_{\rm He}}}$	
	$\frac{v_{\text{He}}}{482} = \sqrt{\frac{32.0}{4.01}}$	
	$\frac{v_{\text{He}}}{482} = \sqrt{8.00} = 2.82$	
	$v_{\rm He} = (2.82)(482) = 1360 \frac{\rm m}{\rm s}$	

Molecular Speed & Effusion

Big Ideas	Details	Unit: Gases
		Homework Problems
	1. Th	e temperature of the air in a room is 25.0 °C. Determine the root mean
	sq th	uare speed of the nitrogen molecules (molar mass 28.0 $rac{g}{mol}$) in the air in at room.
	Ar 2. If t ha m	Inswer: $515 \frac{m}{s}$ the N ₂ molecules (which have a molar mass of $28.0 \frac{g}{mol}$) in a sample of air twe a root mean square speed of $450 \frac{m}{s}$, what is the speed of the O ₂ olecules (which have a molar mass of $32.0 \frac{g}{mol}$)?
	Ar	nswer: 421 <u>m</u>
	3. A ox es	balloon is filled with 1.0 mole each of helium (molar mass 4.0 $\frac{g}{mol}$) and tygen (molar mass 32 $\frac{g}{mol}$). If half of the gas in the balloon is allowed to cape, how many moles each of helium and oxygen are left in the balloon?
	(H sa re ba	lint: Find the ratio of the speeds of the two molecules. This will be the me as the ratio of the number of moles of each gas that escapes. The ciprocal will be the ratio of the number of moles of each gas left in the alloon.)
	Ar	nswer: 0.26 mol He, 0.74 mol O_2

Unit: Atomic Structure

Details

Topics covered in this chapter:

Atomic Structure	154
History of Atomic Theory	157
Conservation of Mass; Definite & Multiple Proportions	164
Fundamental Forces	166
The Standard Model	167
Average Atomic Mass	173

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

HS-PS1-1: Use the periodic table as a model to predict the relative properties of main group elements, including ionization energy and relative sizes of atoms and ions, based on the patterns of electrons in the outermost energy level of each element. Use the patterns of valence electron configurations, core charge, and Coulomb's law to explain and predict general trends in ionization energies, relative sizes of atoms and ions, and reactivity of pure elements.

Atomic Structure

Unit: Atomic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Identify subatomic particles and their locations within the atom.

Success Criteria:

- Protons and neutrons are correctly located in the nucleus.
- Electrons are correctly located outside the nucleus.
- Relative masses of protons, neutrons and electrons are correct.
- Chemical symbols are written correctly with correct value and placement of atomic symbol, atomic number, atomic mass and charge.

Tier 2 Vocabulary: nucleus, charge

Language Objectives:

• Correctly describe the parts of the atom and their locations within the atom.

Notes:

atom: the smallest piece of an element that retains the properties of that element.

- <u>nucleus</u>: a dense region in the center of an atom. The nucleus is made of protons and neutrons, and contains almost all of an atom's mass.
- proton: a subatomic particle found in the nucleus of an atom. It has a charge of +1, and a mass of 1 atomic mass unit (amu).
- <u>neutron</u>: a subatomic particle found in the nucleus of an atom. It has no charge (is neutral), and has a mass of 1 amu.
- <u>electron</u>: a subatomic particle found *outside* the nucleus of an atom. It has charge of -1 and a mass of 0 amu (really about $1/_{2000}$ amu). Atoms can gain, lose, or share electrons in chemical reactions.
- <u>charge</u>: electrical charges can be positive or negative. Opposite cancel each other out, so the charge of an atom is the difference between how many positive charges (protons) it has, and how many negative charges (electrons) it has. For example, a chlorine atom with 17 protons (+17) and 18 electrons (-18) would have a charge of -1. (The difference is 1, and it's negative because it has more negatives than positives.)

<u>neutral atom</u>: an atom with a charge of zero (positives = negatives).

Atomic Structure

Big Ideas	Details	Unit: Atomic Structure
	ion: an atom or molecule that has a charge (has either gained or lost electrons. This (electrons) than positives (protons), or m (electrons).	either positive or negative), because it means it has either more negatives nore positives (protons) than negatives
	atomic number (Z): the identity of an atom i charge in its nucleus. (This works becaus given to or shared with another atom.) T protons in the nucleus. Each element ha	is based on the amount of (positive) se particles from the nucleus cannot be The atomic number is the number of as a unique atomic number.
	mass number (A): the total number of proto atom. (Protons and neutrons each have the electrons are so small that their mass whole number that is closest to the atom	ns + neutrons in the nucleus of an a mass of almost exactly 1 amu, and s is negligible.) Generally equal to the nic mass.
	atomic mass: the actual mass of an atom; th neutrons, and electrons (minus a small a energy to hold the atom together). Alwa	e sum of the masses of its protons, mount of mass that is converted to ays close in value to the mass number.
	isotopes: atoms of the same element (same but that have different numbers of neutr numbers) from each other.	atomic number = same # of protons), rons (and therefore different mass
	Isotopes are described by their mass nun 6 protons and 6 neutrons, which gives it has 6 protons and 8 neutrons, which give	nbers. For example, carbon-12 (¹² C) has a mass number of 12. Carbon-14 (¹⁴ C) es it a mass number of 14.
	element symbol: a one- or two-letter abbrev are given temporary three-letter symbols is always capitalized. Other letters in an <i>This is important to remember</i> . For exam the compound carbon monoxide, which oxygen.	viation for an element. (New elements s.) The first letter in an element symbol element symbol are always lower case. nple, Co is the element cobalt, but CO is contains the elements carbon and

		A	tomic St	ructure			Page: 156
Big Ideas	Details					Unit: Atom	nic Structure
-	<u>chemical symbol</u> : a shorthand notation that shows information about an element, including its element symbol, atomic number, mass number, and charge. For example, the symbol for a magnesium-25 ion with a +2 charge would be:						
	$^{25}_{12}$ Mg ²⁺ This notation shows the element symbol for magnesium (Mg) in the center, th atomic number (12, because it has 12 protons) on the bottom left, the mass number (25, because it has 12 protons + 13 neutrons = 25 amu) on the top lef and the charge (2+, which means it has somehow lost two of its electrons) on the top right. By convention, ions are labeled with the charge number before the charge sign, so we write 2+ instead of +2.						
					center, the ne mass ne top left, trons) on er before		
	Each rov given, tl element	Homework Problems ach row in this table is like a Sudoku puzzle. For each row, use the numbers ven, the relationships between the columns, and the periodic table of the ements to fill in the rest of the row. Use the first row as an example.				numbers of the e.	
	symbol atomic # mass # protons neutrons electrons					charge	
	⁴¹ ₂₀ Ca ²⁺	20	41	20	21	18	+2
	В				6	5	
			56	24			0
	Ca ²⁺				20		
		60			84	57	
			207			80	+2
					0	0	+1
	Kr		84				0

Use this space for summary and/or additional notes:

35

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Unit: Atomic Structure

Details

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Give a timeline for the development of atomic theory.
- Explain how each discovery changed our model of the atom.

Success Criteria:

- Discoveries are in the correct chronological order.
- Descriptions explain how each new discovery added to or changed the model of the atom.

Tier 2 Vocabulary: model,

Language Objectives:

• Correctly describe the parts of the atom and their locations within the atom.

Notes:

<u>atomic theory</u>: a theory that explains behavior of chemical elements based on the atoms that they are made of, and the composition of those atoms.

Modern Atomic Theory

The current model (theory) of the atom is the <u>quantum mechanical model</u>. It states that:

- The atom contains a nucleus at the center. The nucleus contains most of the mass of the atom. The nucleus consists of:
 - protons (positively charged)
 - neutrons (neutral)
- The atom contains electrons (negatively charged) outside the nucleus.
- Electrons can be added to or removed from atoms. An atom that has gained or lost electrons is called an ion.
- Each electron in an atom is confined to one of several specific regions around the nucleus (called orbitals), but each electron may move freely within its orbital. These orbitals are regions, but they do not have solid boundaries; each electron remains within its orbital because of a balance of forces, which are determined by how much energy the electron has.

Big Ideas	, Details	Unit: Atomic Structure
	Historical Development of Atomic Theory	
	Democritus: ancient Greek philosopher. Credited w (~400 B.C.E.). The theory of Democritus held "atoms", which are physically, but not geom between atoms, there lies empty space; that have always been and always will be in mot number of atoms and of kinds of atoms, wh mass.	with the first theory of atoms d that everything is composed of netrically, indivisible; that at atoms are indestructible, and ion; that there is an infinite ich differ in shape, size and
	John Dalton: English chemist, physicist and meteoro theory that described what atoms are and how	blogist. Credited with the first they behave:
	Dalton's Atomic Theory (1807-08):	
	 everything is made of atoms 	
	 atoms of the same element are identical; at different 	oms of different elements are
	 atoms are not created or destroyed in chem reactions are simple rearrangements of the compounds. 	nical reactions. Chemical atoms into different
	 every sample (molecule) of a compound consame proportions ("Law of Constant Compositions") 	ntains the same atoms in the osition")
	 atoms in compounds occur in simple, whole Multiple Proportions") 	e-number ratios ("Law of
	<u>J.J. Thomson</u> : English physicist. Discovered the elect was to apply an electric current to a gas. This cr negatively-charged electric particles, which app (positive electric terminal). Thomson determine named "corpuscles") came from the atoms that discovery was important because it was the first divisible. Thomson received the Nobel Prize for discovery.	ctron (1897). His experiment reated cathode rays—rays of eared to come from the cathode ed that these particles (which he the cathode was made of. This t evidence that atoms were Physics in 1906 for this
	<u>"plum pudding" model</u> : (1904) J.J. Thomson compa plum pudding with raisins. (Plum pudding, whic in England.) The "pudding" was the positively cl the atom was made of, and the "raisins" were tl (which he called "corpuscles"). Thomson publis called <i>The Corpuscular Theory of Matter</i> .	red the atom with a bowl of ch is a lot like oatmeal, is popular harged substance that most of he negatively-charged electrons hed this theory in 1907 in a book

Big Ideas	Details	Unit: Atomic Structure
	planetary model: (early 1900s) The atom was system. In the 1906 physics textbook, A Millikan and Henry Gale credited Thomso was called <i>electron theory</i>):	s compared with a miniature solar <i>First Course in Physics,</i> authors Robert on with this model (which at the time
	"But since the atoms are probably electr assume that they contain equal amounts Since, however, no evidence has yet app charged electrons ever become detache forward the hypothesis that perhaps the nucleus of the atom about the center of rapidly rotating.	ically neutral, it is necessary to s of positive and negative electricity. beared to show that positively d from molecules, Thomson brings e positive charges constitute the which the negative electrons are
	"According to this hypothesis, then, an a system whose members, the electrons, a diameter of the atom than is the earth w earth's orbit. Furthermore, according to of these electrons which give rise to ligh through conductors of electrons which r which constitutes an electric current in a upon a body which constitutes a static n electrons which constitutes a positive ch	atom is a sort of infinitesimal solar are no bigger with respect to the vith respect to the diameter of the o this hypothesis, it is the vibrations t and heat waves; it is the streaming nave become detached from atoms a metal; it is an excess of electrons egative charge, and a deficiency of narge."
	Robert Millikan: American physicist. Measu based on the rate that oil drops fall throu common factor in all of the measuremen charge—the electron. Millikan received this discovery.	red the electrical charge on an electron ugh an electric field (1909). The its must be the basic particle of electric the Nobel Prize for Physics in 1923 for



Big Ideas	Details	Unit: Atomic Structure
	Bohr model: (1913) Danish physicist Niels Bohr hypothesized that electrons moved around the nucleus as in the planetary model, and the distance of each electron from the nucleus was determined by the amount of energy it had. The energy was quantized, so only specific orbits were allowed. These quantum values of energy could be described by a quantum number (<i>n</i>).	$n=3$ $n=2$ $n=1$ $\Delta E=hv$
	Bohr's model gained wide acceptance, because it combined three prominent theories of the spectroscopy and quantum theory.	time: electron theory,
	Even though the Bohr model of the atom has been su mechanical model, the Bohr model is frequently taug middle school science classes because it is easier to vi relates the atom to the solar system, which is already	perseded by the quantum ht today in elementary and sualize and because it familiar.
	The Bohr model is described in more detail in the sect the Hydrogen Atom," which begins on page 199.	tion "The Bohr Model of
	<u>quantum mechanics</u> : in 1900, German physicist Max Plan absorption and emission of energy that produces ligh called "quanta".	ck postulated that t occurs in discrete packets
	<u>photoelectric effect</u> : In 1905, German physicist Albert Ein energy from light could cause electrons to be emitted energy from this light agreed with Planck's equation, minimum amount of energy specific to each metal tha the electrons, and that this energy was quantized—th release the electrons was all-or-nothing. Einstein reco Physics in 1921 for this discovery.	istein discovered that I from a metal, that the that there was a certain at was required to drive off he energy needed to eived the Nobel Prize in

Big Ideas Details Unit: Atomic Structure Louis de Broglie: In 1924, French physicist Louis de Broglie suggested that matter can act as both a particle and a wave. He theorized that the reason that only integer values for quantum numbers were possible was because as the electron orbits the nucleus, its path must be an integer multiple of the wavelength: n = 6 n = 5 n = 3 n = 4 Erwin Schrödinger: Austrian physicist. Expressed de Broglie's hypothesis in mathematical form (the Schrödinger wave equation) and used it to predict the quantum energies of atoms (1926). The solutions to Schrödinger's equation defined additional integer quantum numbers (ℓ and m) that specified the arrangements of electrons within the atom. These solutions supported the idea that an electron is either able to be detected (present), or unable to be detected (absent), as would be the case for a wave that is detectable at an antinode, but not at a node. Schrödinger's equations resulted in maps of regions around the nucleus of an atom (later named "orbitals," based on the probabilities of finding an electron in the different regions as a function of the energy of the electron. Schrödinger received the Nobel Prize for Physics in 1933 for his work. Sir James Chadwick: British physicist. Discovered the neutron (in 1932), which accounted for previously unexplained mass within an atom. His experiment was to collide alpha particles into beryllium, which caused neutral particles with the same mass as a proton to be ejected. Because these particles were neutrally charged, Chadwick named them neutrons. Chadwick received the Nobel Prize for Physics in 1935 for this discovery.

Details **Unit: Atomic Structure Big Ideas** Homework Make a timeline of how the theory of the atom developed, including the models of Democritus, Dalton, the "plum pudding" model, the planetary model, Bohr, de Broglie and Schrödinger. For each entry, your timeline should include: • a sketch of what the atom might have looked like according to the model • the year the model was proposed • the name(s) of the scientist(s) credited with the model • a 1–2 sentence description of the model Here is an example of what the timeline entry for Democritus might look like: ~400 BCE Democritus Everything is made of indestructible atoms of different sizes, shapes and masses. These atoms are in constant motion.

Big Ideas

Conconvotion of	Mass: Dofinito & Multiple Proportions
	viass; Definite & Wuitiple Proportions
Unit: Atomic Structure	
MA Curriculum Framewo	orks (2016): HS-PS1-1
Mastery Objective(s): (S	tudents will be able to)
 Explain the laws of co proportions. 	onservation of mass, definite proportions, and multiple
 Solve problems relati 	ing to the conservation of mass.
Success Criteria:	
 Explanations account 	for observations about the way atoms combine.
 Solutions account for 	all mass before and after some change.
Tier 2 Vocabulary: conse	rvation
Language Objectives:	
• Explain the laws of conception proportions.	onservation of mass, definite proportions, and multiple
Notes:	
conservation of mass: mat changed in form. All of change took place is pr	ter (mass) can neither be created nor destroyed, only the mass that was present before a chemical or physic esent after the change.
This law holds for the to of atom (element).	otal mass, and also individually for the mass of each typ
For example, in the che	emical equation:
	$HCI + NaOH \rightarrow NaCI + H_2O$

- 1. The combined mass of HCl and NaOH before the reaction is equal to the combined mass of NaCl and H_2O produced by the reaction.
- 2. The mass of each element before the reaction is equal to the mass of that same element after. For example, the number of grams of chlorine in the HCl that reacts is equal to the grams of chlorine in the NaCl produced.

Conservation of Mass; Definite & Multiple Proportions Page: 165

Big Ideas	Details Unit: Atomic Structur	e
	Law of Constant Composition (Law of Definite Proportions): the same compound always contains atoms of the same elements in the same proportions by mass. <i>E.g.</i> , water (H ₂ O) always contains 11 % hydrogen and 89 % oxygen by mass.	_
	The Law of Constant Composition was part of Dalton's theory of atoms, first published in 1803.	
	Note also that the reverse is not necessarily true—very different compounds can have the same atoms in the same proportions, and even the exact same chemical formulas. For example, the compounds ethyl acetate and butyric acid both have the same chemical formula ($C_4H_8O_2$). However, ethyl acetate smells like nail polish, whereas butyric acid smells like a combination of rancid butter and vomit.	
	Law of Multiple Proportions: elements always combine in simple, whole-number ratios. (This works whether you're comparing atoms or masses.) For example, copper and chlorine can combine to form CuCl or CuCl ₂ , but they won't combin to form ratios like Cu _{1.7} Cl _{4.83} .	е
	There is a joke whose punch line depends on the law of multiple proportions:	
	A chemist and her friend walk into a bar. The chemist tells the bartender, "I'd like a glass of H_2O , please." Her friend says, "I'd like H_2O too." Both drink, and the friend dies.	
	The basis of the punch line is that " H_2O too" sounds like " H_2O_2 ," which is hydrogen peroxide.	
	The Law of Multiple Proportions was also first proposed by John Dalton in 1803 as part of his theory of atoms.	
	While the chemistry that we will study this year depends on the laws of constant composition and multiple proportions, there are a few unusual compounds whose elemental composition can vary from sample to sample. One example is the iron oxide wüstite, which can contain between 0.83 and 0.95 iron atoms for every oxygen atom, and thus contains anywhere between 23 % and 25 % oxygen.	

Mastery Objective(s): (Students will be able to...)

Fundamental Forces

Unit: Atomic Structure

Details

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Name, describe, and give relative magnitudes of the four fundamental forces of nature.

Success Criteria:

• Descriptions & explanations are accurate and account for observed behavior.

Tier 2 Vocabulary: strong, weak

Language Objectives:

• Describe the fundamental forces, how strong they are and what they act on.

Notes:

All forces in nature ultimately come from one of the following four forces:

<u>strong force</u> (or "strong nuclear force" or "strong interaction"): an attractive force between quarks. The strong force holds the nuclei of atoms together. The energy comes from converting mass to energy.
 Effective range: about the size of the nucleus of an average-size atom.

<u>weak force</u> (or "weak nuclear force" or "weak interaction"): the force that causes protons and/or neutrons in the nucleus to become unstable and leads to beta nuclear decay. This happens because the weak force causes an up or down quark to change its flavor. (This process is described in more detail in the section on The Standard Model of particle physics, starting on page 167) **Strength**: 10^{-6} to 10^{-7} times the strength of the strong force. **Effective range**: about 1/3 the diameter of an average nucleus.

<u>electromagnetic force</u>: the force between electrical charges. If the charges are the same ("like charges")—both positive or both negative—the particles repel each other. If the charges are different ("opposite charges")—one positive and one negative—the particles attract each other.

Strength: about $1/_{137}$ as strong as the strong force.

Effective range: ∞ , but gets smaller as (distance)².

<u>gravitational force</u>: the force that causes masses to attract each other. Usually only observable if one of the masses is very large (like a planet). **Strength**: only 10^{-39} times as strong as the strong force. **Effective range**: ∞ , but gets smaller as (distance)².

The Standard Model

Unit: Atomic Structure

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Describe the fundamental particles that make up atoms and other matter and non-matter.

Success Criteria:

• Explanations correctly describe aspects of the standard model.

Tier 2 Vocabulary: standard, color

Language Objectives:

• Describe particles of the standard model and their properties.

Notes:

The Standard Model is a theory of particle physics that:

- identifies the particles that matter is ultimately comprised of
- describes properties of these particles, including their mass, charge, and spin
- describes interactions between these particles

The Standard Model dates to the mid-1970s, when the existence of quarks was first experimentally confirmed. Physicists are still discovering new particles and relationships between particles, so the model and the ways it is represented are evolving, much like atomic theory and the Periodic Table of the Elements was evolving at the turn of the twentieth century. The table and the model described in these notes represent our understanding, as of 2018. By the middle of this century, the Standard Model may evolve to a form that is substantially different from the way we represent it today.

The Standard Model in its present form does not incorporate dark matter, dark energy, or gravitational attraction.

The Standard Model

 Details
 Unit: Atomic Structu

 The Standard Model is often presented in a table, with rows, columns, and colorcoded sections used to group subsets of particles according to their properties.

As of 2018, the standard model is represented by a table similar to this one:





Quarks

Quarks are particles that participate in strong interactions (sometimes called the "strong force") through the action of "color charge" (which will be described later). Because protons and neutrons (which make up most of the mass of an atom) are made of three quarks each, quarks are the subatomic particles that make up most of the ordinary matter in the universe. (Dark matter, which accounts for 84.5% of the total matter in the universe, is made from other types of subatomic particles.)

- quarks have color charge (*i.e.*, they interact via the strong force)
- quarks have spin of $+\frac{1}{2}$
- "up-type" quarks have a charge of $+\frac{1}{2}$; "down-type" quarks have a charge of $+\frac{1}{2}$.

Use this space for summary and/or additional notes:

Big Ideas

	The Standard Model Page: 10	59
Big Ideas	Details Unit: Atomic Structu	re
	There are six flavors [*] of quarks: up and down, charm and strange, and top and bottom. Originally, top and bottom quarks were called truth and beauty.	
	Leptons	
	Leptons are the smaller particles that make up most matter. The most familiar lepton is the electron. Leptons participate in "electroweak" interactions, meaning combinations of the electromagnetic and weak forces.	
	 leptons do not have color charge (<i>i.e.</i>, they do not interact via the strong force) 	
	• leptons have spins of $+\frac{1}{2}$	
	 electron-type leptons have a charge of −1; neutrinos do not have a charge. 	
	 neutrinos oscillate, which makes their mass indefinite. 	
	Gauge Bosons	
	Gauge bosons are the particles that carry force—their interactions are responsible for the fundamental forces of nature: the strong force, the weak force, the electromagnetic force and the gravitational force. The hypothetical particle responsible for the gravitational force is the graviton, which has not yet been detected (as of 2018).	
	 photons are responsible for the electromagnetic force. 	
	 gluons are responsible for the strong interaction (strong force) 	
	 W and Z bosons are responsible for the weak interaction (weak force) 	
	Scalar Bosons	
	At present, the only scalar boson we know of is the Higgs boson, which is responsible for mass.	
	* Yes, "flavors" really is the correct term. Blame the 1960s.	

The Standard Model



Fermions

Quarks and leptons are fermions. Fermions are described by Fermi-Dirac statistics and obey the Pauli exclusion principle (which states that no two particles in an atom may have the same exact set of quantum numbers—which are numbers that describe the energy states of the particle).

Fermions are the building blocks of matter. They have a spin of ½, and each fermion has its own antiparticle. (The antiparticle of a fermion is identical to its corresponding particle, but has a charge of the opposite sign. Antiparticles have the same name as the corresponding particle with the prefix "anti-"; for example, the antiparticle of a tau neutrino is a tau antineutrino. Note, however, that for historical reasons an antielectron is usually called a positron.)

Big Ideas	Details	Unit: Atomic Structure
	Bosons	
	Bosons are described by Bose-Einstein statistics, have intege the Pauli Exclusion Principle. Interactions between boson ar and mass.	r spins and do not obey e responsible for forces
	Each of the fundamental bosons is its own antiparticle, excep (whose antiparticle is the W ⁺ boson).	ot for the W⁻ boson
	Hadrons	
	Hadrons are a special class of strongly-interacting composite they are comprised of multiple individual particles). Hadrons fermions. Hadrons composed of strongly-interacting fermion hadrons composed of strongly-interacting bosons are called	particles (meaning that s can be bosons or ns are called baryons; mesons.
	Baryons	
	The most well-known baryons are protons and neutrons, wh three quarks. Protons are made of two up quarks and one d carry a charge of +1. Neutrons are made of one up quark an ("udd"), and carry a charge of zero.	ich each comprised of own quark ("uud"), and d two down quarks
	Some of the better-known baryons include:	
	 nucleons (protons & neutrons). 	
	 hyperons, <i>e.g.</i>, the Λ, Σ, Ξ, and Ω particles. These contiquarks, and are much heavier than nucleons. 	ain one or more strange
	 various charmed and bottom baryons. 	
	 pentaquarks, which contain four quarks and an antiquation 	ark.
	Mesons	
	Ordinary mesons are comprised of a quark plus an antiquark pion, kaon, and the J/ Ψ . Mesons mediate the residual strong nucleons.	. Examples include the g force between
	Some of the exotic mesons include:	
	 tetraquarks, which contain two quarks and two antiquing 	arks.
	 glueball, a bound set of gluons with no quarks. 	
	 hybrid mesons, which contain one or more quark/antion more gluons. 	quark pairs and one or

Big Ideas	Details			Unit: Atomic	c Structure
	Color Charge				
	Color charge is the property that is responsible for the strong nuclear interaction All electrons and fermions (particles that have half-integer spin quantum number must obey the Pauli Exclusion Principle, which states that no two particles with the same larger particle (such as a hadron or atom) can have identical sets of quantum numbers. For electrons, (as you learned in chemistry), if two electrons share the same orbital, they need to have opposite spins. In the case of quarks quarks have a spin of $+\frac{1}{2}$, so in order to satisfy the Pauli Exclusion Principle, if proton or neutron contains three quarks, there has to be some other quantum property that has different values for each of those quarks. This property is ca "color charge" (or sometimes just "color [*] ").			raction. numbers) within of ctrons uarks, all ule, if a ntum is called	
	The "color" ((named afte subatomic p "colorless". white light, w Quarks can e	property has three w r the primary colors article, the colors ha (Recall that combin which is colorless.)	values, which are calle of light). When there ave to be different, ar ing each of the prima ge by emitting a gluor	ed "red," "green," and e are three quarks in nd have to add up to ry colors of light proc n that contains one co	d "blue" a duces blor and
	one anticolo change. For	r. Another quark at example, suppose a	osorbs the gluon, and a blue quark emits a b	both quarks undergo lue antigreen gluon:	o color
		0		0	
		\bigcirc \bigcirc	ÓO	00	
	You can imag antigreen" g became a gr quark and ca receiving qu one red, one	gine that the quark luon). Because it al een quark. Meanwl ancels its color. The ark to become blue. green, and one blu	sent away its own blu so sent out antigreen hile, the antigreen par blue from the blue ar After the interaction e quark, which means	e color (the "blue" in , it was left with gree rt of the gluon finds t ntigreen gluon causes n, the particle once ag s color charge is cons	the "blue n so it he green s the gain has erved.
	[*] Just like "spi spinning, "c couldn't pos larger than	n" is the name of a pr olor" is a property tha ssibly have actual colo quarks!	operty of energy that has nothing to do with thas nothing to do with or—the wavelengths of v	as nothing to do with a n actual color. In fact, c visible light are thousar	ctual quarks ids of times
	Use this space	ce for summary and	/or additional notes:		

Unit: Atomic Structure

Details

MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to...)

• Calculate the average atomic mass of an atom from percent abundance data.

Success Criteria:

- Solutions correctly turn masses into percentages.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: abundance

Language Objectives:

• Explain the laws of conservation of mass, definite proportions, and multiple proportions.

Notes:

<u>mass number</u>: the mass of *one individual atom* (protons + neutrons). Always a whole number.

abundance: the percentage of atoms of an element that are one specific isotope.

<u>average atomic mass</u>: the estimated weighted *average* of the mass numbers *of all of the atoms* of a particular element on Earth.

Analogy: average atomic mass works the same way as class average on a test.

- 1. Multiply each score times the number of students who got it.
- 2. Add up the number for each score to get the total points.
- 3. Divide the total by the number of students to get class average.

Details Unit: Atomic Structure			
Problem:			
The atomic mass and abundance of the two stable isotopes of carbon are:			
Isotope	Atomic Mass (amu)	Relative Abundance	
carbon-12	12.000 000	98.93 %	
carbon-13	13.003 355	1.07 %	
What is the average ator	nic mass of carbon?		
How to solve:			
 Convert percent a Multiply the fract Add up the sub-to Check that your a 	abundances to fractions ional abundance times otal from each isotope to verage atomic mass is in	(divide by 100). the atomic mass for each isotope. o get the total atomic mass. n between the smallest and largest.	
Answer:			
1. Convert abundan	ces to fractions		
98.93 % ÷ 10	00 = 0.9893 1.07	% ÷ 100 = 0.0107	
2. Multiply abundar 0.9893 x 12.000 0	ce x mass # for each iso 00 = 11.8716	tope	
0.0107 X 13.003 3	55 = 0.1391		
3. Add up the numb	er from each isotope to	get the total	
11.8716 + 0.1391	= 12.0107		
4. Check that your a isotope and the n Yes, 12.0107 is be	nswer is in between the nass number of the larg etween 12 and 13.	e mass number of the smallest est one.	
	Details Problem: The atomic mass and abuters Isotope carbon-12 carbon-13 What is the average atom How to solve: 1. Convert percent at 2. Multiply the fract 3. Add up the sub-tom 4. Check that your at 98.93 % ÷ 10 2. Multiply abundant 0.9893 x 12.0000 0.0107 x 13.003 3 3. Add up the numb 11.8716 + 0.1391 4. Check that your at isotope and the m Yes, 12.0107 is bether	DetailsProblem:The atomic mass and abundance of the two stateIsotopeAtomic Mass (amu)carbon-1212.000 000carbon-1313.003 355What is the average atomic mass of carbon?How to solve:1. Convert percent abundances to fractions2. Multiply the fractional abundance times to3. Add up the sub-total from each isotope to4. Check that your average atomic mass is inAnswer:1. Convert abundances to fractions98.93 % ± 100 = 0.98931.072. Multiply abundance x mass # for each iso0.9893 x 12.000 000 = 11.87160.0107 x 13.003 355 = 0.13913. Add up the number from each isotope to11.8716 + 0.1391 = 12.01074. Check that your answer is in between the isotope and the mass number of the larger Yes, 12.0107 is between 12 and 13.	

Big Ideas Details **Homework Problems** Calculate the average atomic mass of each of the following elements, based on the percent abundance of their isotopes. For each element, your answers should agree with the atomic mass listed on the periodic table. Because you can look up the answers, you must show how to set up the calculations in order to receive credit. 1. bromine atomic mass relative isotope (amu) abundance ⁷⁹₃₅Br 78.9184 50.69% $^{81}_{35}{\rm Br}$ 80.9163 49.31% 2. boron atomic mass relative isotope abundance (amu) ¹⁰₅B 10.0129 19.9% ¹¹₅B 11.0093 80.1%

Big Ideas	Details	•		Unit: Atomic S
	3. chlorine			
	isotope	atomic mass (amu)	relative abundance	_
	³⁵ ₁₇ Cl	34.9689	75.78%	
	³⁷ ₁₇ Cl	36.9659	24.22%	
	4 magnesium	2		
	4. magnesium	atomic mass	relative	
	isotope	(amu)	abundance	
	²⁴ ₁₂ Mg	23.9850	78.99%	
	²⁵ ₁₂ Mg	24.9858	10.00 %	
	²⁶ ₁₂ Mg	25.9826	11.01%	
		Ī	<u>1</u>	

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Unit: Nuclear Chemistry

Details

Topics covered in this chapter:

Radioactive Decay1	L 7 8
Nuclear Equations1	83
Mass Defect & Binding Energy1	186
Half-Life1	188
Nuclear Fission & Fusion1	191
Practical Uses for Nuclear Radiation1	194

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

The Nuclear Chemistry unit is not included in the NGSS standards for high school chemistry. Massachusetts moved the topic to the high school physics course, which means this unit may be omitted from a chemistry course that is intended to follow the 2016 Massachusetts Curriculum Frameworks.

Radioactive Decay

Unit: Nuclear Chemistry

MA Curriculum Frameworks (2016): N/A (HS-PS1-8 in physics frameworks) **Mastery Objective(s):** (Students will be able to...)

- Determine the products of alpha, beta-minus, and beta-plus radioactive decay and electron capture.
- Predict the most likely form of radioactive decay for an isotope based on its position relative to the band of stability on a proton-neutron graph.

Success Criteria:

- Form of radioactive decay correctly identified.
- Products of decay correctly identified.
- Correct nuclear equation.

Tier 2 Vocabulary: decay

Language Objectives:

- Explain the processes of alpha, beta-plus and beta-minus radioactive decay and electron capture.
- Understand and correctly use the terms "radioactive decay," "nuclear instability," "alpha decay," "beta decay," "gamma rays," and "penetrating power."

Notes:

<u>nuclear instability</u>: When something is unstable, it is likely to change. If the nucleus of an atom is unstable, changes can occur that affect the number of protons and neutrons in the atom.

Note that when this happens, the nucleus ends up with a different number of protons. This causes the atom to literally turn into an atom of a different element. When this happens, the physical and chemical properties instantaneously change into the properties of the new element!

<u>radioactive decay</u>: the process by which the nucleus of an atom changes, transforming the element into a different element or isotope.

<u>nuclear equation</u>: an equation describing (through chemical symbols) what happens to an atom as it undergoes radioactive decay.

Big IdeasDetailsUnit: Nuclear ChemistryCauses of Nuclear InstabilityTwo of the causes of nuclear instability are:Sizebecause the strong force acts over a limited distance, when nuclei get too large (more than 82 protons), it is no longer possible for the strong force to keep the nucleus together indefinitely. The form of decay that results from an atom exceeding its stable size is called alpha (a) decay.The Weak Nuclear ForceThe weak force is caused by the exchange (absorption and/or emission) of W and Z bosons. This causes a down quark to change to an up quark or vice-versa. The change of quark flavor has the effect of changing a proton to a neutron, or a neutron to a proton. (Note that the action of the weak force is the only known way of changing the flavor of a quark.) The form of decay that results from the action of the weak force is called beta (6) decay.band of stability: isotopes with a ratio of protons to neutrons that results in a stable nucleus (one that does not spontaneously undergo radioactive decay). This observation suggests that the ratio of up to down quarks within the nucleus is somehow involved in preventing the weak force from causing quarks to change flavor.130Image: proton size of the strong flavor.130Image: protong flavor.		Nacioactive Decay Page. 179
Causes of Nuclear Instability Two of the causes of nuclear instability are: Size because the strong force acts over a limited distance, when nuclei get too large (more than 82 protons), it is no longer possible for the strong force to keep the nucleus together indefinitely. The form of decay that results from an atom exceeding its stable size is called alpha (α) decay. The Weak Nuclear Force The weak force is caused by the exchange (absorption and/or emission) of W and Z bosons. This causes a down quark to change a proton to a neutron, or a neutron to a proton. (Note that the action of the weak force is the only known way of changing the flavor of a quark.) The form of decay that results from the action of the weak force is called beta (θ) decay. band of stability : isotopes with a ratio of protons to neutrons that results in a stable nucleus (one that does not spontaneously undergo radioactive decay). This observation suggests that the ratio of up to down quarks within the nucleus is somehow involved in preventing the weak force from causing quarks to change flavor. to much mass; α decay likely $\beta = \frac{1300}{100}$ $\beta = \frac{100}{100}$ $\beta $	Big Ideas	Details Unit: Nuclear Chemistry
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Size because the strong force acts over a limited distance, when nuclei get too large (more than 82 protons), it is no longer possible for the strong force to keep the nucleus together indefinitely. The form of decay that results from an atom exceeding its stable size is called alpha (<i>a</i>) decay. The Weak Nuclear Force The weak force is caused by the exchange (absorption and/or emission) of W and Z bosons. This causes a down quark to change to an up quark or vice-versa. The change of quark flavor has the effect of changing a proton to a neutron, or a neutron to a proton. (Note that the action of the weak force is the only known way of changing the flavor of a quark.) The form of decay that results from the action of the weak force is called beta (<i>θ</i>) decay. <u>band of stability</u> : isotopes with a ratio of protons to neutrons that results in a stable nucleus (one that does not spontaneously undergo radioactive decay). This observation suggests that the ratio of up to down quarks within the nucleus is somehow involved in preventing the weak force from causing quarks to change flavor.		Two of the causes of nuclear instability are:
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band of stability: isotopes with a ratio of protons to neutrons that results in a stable nucleus (one that does not spontaneously undergo radioactive decay). This observation suggests that the ratio of up to down quarks within the nucleus is somehow involved in preventing the weak force from causing quarks to change flavor. too much mass; a decay likely α decay likely α decay likely α decay and α decay likely β decay and β decay and β decay and β decay and β decay of β decay and β decay likely α decay likely		The weak force is caused by the exchange (absorption and/or emission) of W and Z bosons. This causes a down quark to change to an up quark or vice-versa. The change of quark flavor has the effect of changing a proton to a neutron, or a neutron to a proton. (Note that the action of the weak force is the only known way of changing the flavor of a quark.) The form of decay that results from the action of the weak force is called beta (β) decay.
too much mass; α decay likely 130 120 110 α decay likely 100 excess neutrons; β - excess protons; decay likely β + decay and electron capture likely 10 10 10 10 10 10 10 10 10 10		<u>band of stability</u> : isotopes with a ratio of protons to neutrons that results in a stable nucleus (one that does not spontaneously undergo radioactive decay). This observation suggests that the ratio of up to down quarks within the nucleus is somehow involved in preventing the weak force from causing quarks to change flavor.
$\begin{array}{c} 130 \\ 120 \\ 100 \\$		too much mass;
120 - 200 + 200		130– α decay likely
10 20 30 40 50 60 70 80 90 100 Atomic Number, Z		120 - 120 - 110 - excess $110 - excess$ $100 - e$
		Atomic Number, Z

Use this space for summary and/or additional notes:

Radioactive Decay

Big Ideas	Details	Unit: Nuclear Chemistry
	alpha (α) decay: a type of radioactive and two neutrons (an alpha partic nucleus of a helium-4 atom), with	decay in which the nucleus loses two protons cle). An alpha particle is a ${}_{2}^{4}He^{2+}$ ion (the two protons, a mass of 4 amu, and a charge of
	+2. For example: $^{238}_{92}$ U-	$\rightarrow \frac{^{234}}{^{90}}\text{Th} + \frac{^{4}}{^{2}}\text{He}$
	Atoms are most likely to undergo proton/neutron ratio but a large a	alpha decay if they have an otherwise stable atomic number.
	Alpha decay has never been obse 52 (tellurium), and is rare in eleme (tantalum).	rved in atoms with an atomic number less than ents with an atomic number less than 73
	<u>Net effects of α decay:</u>	
	 Atom loses 2 protons and 2 mass number goes down by 	neutrons (atomic number goes down by 2 and 4)
	• An α particle (a ${}^{4}_{2}\text{He}^{+2}$ ion)	is ejected from the nucleus at high speed.
	beta minus (β –) decay: a type of radio	pactive decay in which a neutron is converted
	to a proton and the nucleus ejects	s a high speed electron $({}^{0}_{-1}e$).
	Note that a neutron consists of or proton consists of two up quarks a occurs, the weak force causes one up, which causes the neutron (uu proton was gained, the atomic nu proton used to be a neutron, the	The up quark and two down quarks (udd), and a and one down quark (uud). When β - decay e of the quarks changes its flavor from down to d) to change into a proton (udd). Because a mber increases by one. However, because the mass number does not change. For example:
	3	${}^{2}_{5}P \rightarrow {}^{32}_{16}S + {}^{0}_{-1}e$
	Atoms are likely to undergo β - de enough protons to achieve a stab that are heavier than isotopes of t (because of the "extra" neutrons)	cay if they have too many neutrons and not le neutron/proton ratio. Almost all isotopes the same element within the band of stability undergo β - decay.
	<u>Net effects of β– decay:</u>	
	 Atom loses 1 neutron and gains number does not change) 	1 proton (atomic number goes up by 1; mass
	• A β – particle (an electron) is eje	cted from the nucleus at high speed.
	Note that a β - particle is assigned an electron is some sort of "anti-p equation for the number of proto	an atomic number of -1. <i>This does not mean roton</i> ". The -1 is just used to make the ns work out in the nuclear equation.
Radioactive Decay

ig ideas	Details Unit: Nuclear Chemistry
	beta plus (β +) decay: a type of radioactive decay in which a proton is converted to a neutron and the nucleus ejects a high speed antielectron (positron, $_{+1}^{0}e$).
	With respect to the quarks, β + decay is the opposite of β - decay When β + decay occurs, one of the quarks changes its flavor from up to down, which changes the proton (uud) into a neutron (udd). Because a proton was lost, the atomic number decreases by one. However, because the neutron used to be a proton, the mass number does not change. For example:
	$^{23}_{12}$ Mg $\rightarrow ^{23}_{11}$ Na + $^{0}_{+1}$ e
	Atoms are likely to undergo β + decay if they have too many protons and not enough neutrons to achieve a stable neutron/proton ratio. Almost all isotopes that are lighter than the isotopes of the same element that fall within the band of stability ("not enough neutrons") undergo β + decay.
	Net effects of β + decay:
	 Atom loses 1 proton and gains 1 neutron (atomic number goes down by 1; mass number does not change)
	 A β+ particle (an antielectron or positron) is ejected from the nucleus at high speed.
	electron capture (sometimes called "K-capture"): when the nucleus of the atom "captures" an electron from the innermost shell (the K-shell) and incorporates it into the nucleus. This process is exactly the reverse of β - decay; during electron capture, a quark changes flavor from up to down, which changes a proton (uud) into a neutron (udd):
	$^{23}_{12}$ Mg + $^{0}_{-1}e \rightarrow ^{23}_{11}$ Na
	Note that β + decay and electron capture produce the same products. Electron capture can sometimes (but not often) occur without β + decay. However, β + decay is <u>always</u> accompanied by electron capture.
	Atoms are likely to undergo electron capture (and usually also β + decay) if they have too many protons and not enough neutrons to achieve a stable neutron/proton ratio. Almost all isotopes that are lighter than the isotopes of the same element that fall within the band of stability undergo electron capture, and usually also β + decay.
	Net effects of electron capture:
	 An electron is absorbed by the nucleus.
	• Atom loses 1 proton and gains 1 neutron (atomic number goes down by 1;

Radioactive Decay

Big Ideas	Details	,	Unit: Nuclear Chemistry			
	 gamma (Υ) rays: most radioactive emitted in the form of gammelectromagnetic radiation. (types of electromagnetic radiation, types of electromagnetic radiation, they can pendamage. Because gamma radiate the composition All of the types of radioactive rays. This means to be comparadioactive decay equations 	most radioactive decay produces energy. Some of that energy is ne form of gamma rays, which are very high energy photons of netic radiation. (Radio waves, visible light, and X-rays are other ctromagnetic radiation.) Because gamma rays are waves (which ss), they can penetrate far into substances and can do a lot of ccause gamma rays are not particles, emission of gamma rays does the composition of the nucleus. Dees of radioactive decay mentioned in these notes also produce γ means to be complete, we would add gamma radiation to each of t decay equations described above:				
	$^{238}_{92}U \rightarrow ^{234}_{90}Th + ^{4}_{2}He +$	⁰ ₀ γ ³² ₁₅ P	$\gamma \rightarrow \frac{32}{16}S + \frac{0}{-1}e + \frac{0}{0}\gamma$			
	$^{23}_{12}Mg \rightarrow ^{23}_{11}Na + ^{0}_{+1}e +$	⁰ ₀ γ ²³ ₁₂ Μg	$g + {}^{0}_{-1}e \rightarrow {}^{23}_{11}Na + {}^{0}_{0}\gamma$			
	<u>penetrating power</u> : the distance into/through another substa emission (faster = more pene emission (heaver = less pene	that radioactive particles nce is directly related to etrating) and inversely rel trating):	s can penetrate the velocity of the lated to the mass of the			
	gamma rays	beta particles	alpha particles			
	lightest		heaviest			
	most penetrating power		 least penetrating power 			
	α 🐩 β • γ	aluminium	lead			
	Note also that denser substa absorbing radioactive emissi for experiments involving rad	nces (such as lead) do a b ons. This is why lead is co dioactive substances.	petter job of blocking and ommonly used as shielding			

Nuclear Equations

Unit: Nuclear Chemistry

MA Curriculum Frameworks (2016): N/A (HS-PS1-8 in physics frameworks) **Mastery Objective(s):** (Students will be able to...)

• Write & solve nuclear equations.

Success Criteria:

- Equations include the correct product of decay (α , β or β + particle)
- Equations include the correct starting material(s) and/or product(s).

Tier 2 Vocabulary: decay

Language Objectives:

• Explain the equations for radioactive decay and how to calculate the products.

Notes:

<u>nuclear equation</u>: a chemical equation describing the process of an isotope undergoing radioactive decay. For example:

 $^{238}_{92}U \rightarrow ^{234}_{90}Th + ^{4}_{2}He$

In a nuclear equation, the number of protons (atomic number) and the total mass (mass number) are conserved on both sides of the arrow. If you look at the bottom (atomic) numbers, and replace the arrow with an = sign, you would have the following:

92 = 90 + 2

Similarly, if you look at the top (mass) numbers, and replace the arrow with an = sign, you would have:

238 = 234 + 4

Nuclear Equations

Big Ideas	Details	Unit: Nuclear Chemistry
	Sample problems:	
	Q: What are the products of beta-minus (β –) decay of ¹³¹ I?	
	A: A β - particle is an electron, which we write as $\int_{-1}^{0} e$ in a r	nuclear equation. This
	means ¹³¹ I decays into some unknown particle plus $^{0}_{-1}e$. The equation is:
	${}^{131}_{53}\mathrm{I} \rightarrow {}^m_p X + {}^0_{-1} e$	
	We can write the following equations for the atomic and	d mass numbers:
	Atomic #s: $53 = p + -1 \rightarrow p = 54$; therefore X is Xe	
	Mass #s: $131 = m + 0 \rightarrow m = 131$	
	Therefore, particle X is $^{131}_{54}$ Xe So our final answer is:	
	The two products of decay in this reaction are $\frac{131}{54}$ Xe and	d $_{-1}^{0}e$.
	Q: Which particle was produced in the following radioactiv	e decay reaction:
	${}^{212}_{86}$ Rn $\rightarrow {}^{208}_{84}$ Po + ${}^{m}_{\rho}X$	
	A: The two equations are:	
	Atomic #s: $86 = 84 + p \rightarrow p = 2$; therefore X is He	
	Mass #s: $212 = 208 + m \rightarrow m = 4$	
	Therefore, particle X is ${}_{2}^{4}$ He , which means it is an $lpha$ part	ticle.

Use this space for summary and/or additional notes:

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Nuclear Equations

Big Ideas	Details		Unit: Nuclear Chemistry
		Homework Problems	
	For these problems, you information from "Table Reference Tables.	u will need to use a periodic table and e use a periodic table and e U. Selected Radioisotopes" on particular particular set of the periodic table and the periodic table and ta	nd radioactive decay ge 514 of your Chemistry
	Give the nuclear equation	on(s) for radioactive decay of the fo	llowing:
	1. ²²² Rn		
	2. ⁸⁵ Kr		
	3. ²²⁰ Fr		
	4. ³⁷ K		
	5. ³ H		
	Give the starting mater	ial for the following materials produ	used by radioactive decay:
		reculting in ²⁶⁷ Hc	iced by fauloactive decay.
	7. Beta-minus (β−)	decay resulting in ¹⁸⁵ ₇₅ Re	

Use this space for summary and/or additional notes:

Unit: Nuclear Chemistry

Details

MA Curriculum Frameworks (2016): N/A (HS-PS1-8 in physics frameworks) **Mastery Objective(s):** (Students will be able to...)

- Explain where the energy that powers the strong force comes from.
- Calculate the "missing" mass and convert it to energy.

Success Criteria:

- Theoretical mass accounts for mass of all protons, neutrons and electrons.
- Mass defect is calculated correctly.
- Einstein's equation (*E* = *mc*²) is properly applied to the mass defect to calculate energy.

Tier 2 Vocabulary: defect

Language Objectives:

• Explain the concept of "missing" mass and how it converts to energy.

Notes:

<u>binding energy</u>: the energy that holds the nucleus of an atom together through the strong nuclear force

The binding energy comes from the small amount of mass (the mass defect) that was turned into a large amount of energy, given by the equation:

 $E = mc^2$

where *E* is the binding energy, *m* is the mass defect, and *c* is the speed of light $(3 \times 10^8 \frac{\text{m}}{\text{c}})$, which means c^2 is a very large number).

<u>mass defect</u>: the difference between the actual mass of an atom, and the sum of the masses of the protons, neutrons, and electrons that it contains.

- A proton has a mass of 1.6726×10^{-27} kg = 1.0073 amu
- A neutron has a mass of 1.6749×10^{-27} kg = 1.0087 amu
- An electron has a mass of 9.1094×10^{-31} kg = 0.0005486 amu

To calculate the mass defect, total up the masses of each of the protons, neutrons, and electrons in an atom. The actual (observed) atomic mass of the atom is always *less* than this number. The "missing mass" is called the mass defect.

Mass Defect & Binding Energy

Big Ideas	Details Unit: Nuclear Chemistry
	Sample problem:
	Q: Find the mass defect of 1 mole of uranium-238.
	A: $^{238}_{92}$ U has 92 protons, 146 neutrons, and 92 electrons. This means the total mass
	of one atom of $^{238}_{92}$ U should theoretically be:
	92 protons × 1.0073 amu = 92.6704 amu
	146 neutrons × 1.0087 amu = 147.2661 amu
	92 electrons × 0.000 5486 amu = 0.0505 amu
	92.6704 + 147.2661 + 0.0505 = 239.9870 amu
	The actual observed mass of one atom of $^{238}_{92}$ U is 238.0003 amu.
	The mass defect is therefore 239.9870 – 238.0003 = 1.9867 amu.
	One mole of $^{238}_{92}$ U would have a mass of 238.0003 g, and therefore a total mass defect of 1.9867 g (which is 0.0019867 kg).
	Because E = mc ² , that means the binding energy of one mole of $^{238}_{92}$ U is:
	$0.0019867 \text{kg} \times (3.00 \times 10^8)^2 = 1.79 \times 10^{14} \text{J}$
	In case you don't realize just how large that number is, the binding energy of just 238 g (1 mole) of $^{238}_{92}$ U would be enough energy to heat every house on Earth for an entire winter!

Half-Life

Unit: Nuclear Chemistry

Details

Big Ideas

MA Curriculum Frameworks (2016): N/A (HS-PS1-8 in physics frameworks) **Mastery Objective(s):** (Students will be able to...)

- Determine the amount of radioactive material remaining after an integer number of half-lives.
- Determine the amount of time that has elapsed based on the fraction of radioactive material remaining (*e.g.*, carbon dating).

Success Criteria:

- Solutions use the appropriate equation for the information given.
- Solutions have the correct quantities substituted for the correct variables.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: half-life, decay

Language Objectives:

• Explain how exponential decay works.

Notes:

The atoms of radioactive elements are unstable, and they spontaneously decay (change) into atoms of other elements.

For any given atom, there is a certain probability, P, that it will undergo radioactive decay in a given amount of time. The half-life, τ , is how much time it would take to have a 50 % probability of the atom decaying. If you start with n atoms, after one half-life, half of them (0.5n) will have decayed.

Amount of Material Remaining

If we start with 32 g of 53 Fe, which has a half-life (τ) of 8.5 minutes, we would observe the following:

#	# minutes	0	8.5	17	25.5	34
‡	# half lives	0	1	2	3	4
ā	amount left	32 g	16 g	8 g	4 g	2 g

Half-Life

Big Ideas	Det	tails					I	Unit: Nu	clear Chemistry
	Finding the Time that has Passed (integer number of half-lives)								
	If the amount you started with divided by the amount left is an exact power of two,								
	γοι	ı have an int	eger number o	f half-liv	ves and	you can	make a	table.	
	Sa	mple prob	lem:						
	Q: If you started with 64 g of ¹³¹ I, how long would it take until there was only 4 g remaining? The half-life (τ) of ¹³¹ I is 8.07 days.								
	A: $\frac{64}{4} = 16$ which is a power of 2, so we can simply make a table:								
			# half lives	0	1	2	3	4	
			amount remaining	64 g	32 g	16 g	8 g	4 g	
		From the ta	ble, after 4 ha	f-lives,	we have	4 g rem	aining.		
	The half-life ($ au$) of ¹³¹ I is 8.07 days.								
	8.07 × 4 = 32.3 days								
	Finding the amount remaining and time that has passed for a non-integer number of half-lives requires logarithms, and is beyond the scope of this course.								
	Homework Problems								
	For these problems, you will need to use half-life information from "Table U. Selected Radioisotopes" on page 514 of your Chemistry Reference Tables.								
	 If a lab had 128 g of ³H waste 49 years ago, how much of it would be left today? (Note: you may round off to a whole number of half-lives.) 					ould be left <i>lives.</i>)			
		Answer	:8g						

Big Ideas	Details	Unit: Nuclear Chemistry
	2.	Suppose a student stole a 20. g sample of ⁴² K at 8:30am on Friday. When the student was called down to the vice principal's office on Monday at the convenient time of 10:54am, how much of the ⁴² K was left?
		Answer: 0.31 g
	3.	If a school wants to dispose of small amounts of radioactive waste, they can store the materials for ten half-lives, and then dispose of the materials as regular trash.
		a. If we had a sample of ³² P, how long would we need to store it before disposing of it?
		 Answer: 143 days b. If we had started with 64 g of ³²P, how much ³²P would be left after ten half-lives? Approximately what fraction of the original amount would be left?
	1	Answer: 0.063 g; approximately $\frac{1}{1000}$ of the original amount.
	4.	expected amount of ¹⁴ C, how old is the sample?
		(Hint: pretend you started with 1 g of $^{14}\mathrm{C}$ and you have 0.25 g remaining.)
		Answer: 11 460 years

Nuclear Fission & Fusion

Unit: Nuclear Chemistry

Details

MA Curriculum Frameworks (2016): N/A (HS-PS1-8 in physics frameworks) **Mastery Objective(s):** (Students will be able to...)

- Identify nuclear processes as "fission" or "fusion".
- Describe the basic construction and operation of fission-based and fusion-based nuclear reactors.

Success Criteria:

• Descriptions account for how the energy is produced and how the radiation is contained.

Tier 2 Vocabulary: fusion, nuclear

Language Objectives:

• Explain how fission-based and fusion-based nuclear reactors work.

Notes:

Fission

<u>fission</u>: splitting of the nucleus of an atom, usually by bombarding it with a highspeed neutron.

When atoms are split by bombardment with neutrons, they can divide in hundreds of ways. For example, when ²³⁵U is hit by a neutron, it can split more than 200 ways. Three examples that have been observed are:

Note that each of these bombardments produces more neutrons. A reaction that produces more fuel (in this case, neutrons) than it consumes will accelerate. This self-propagation is called a <u>chain reaction</u>.

Note also that the neutron/proton ratio of ²³⁵U is about 1.5. The stable neutron/proton ratio of each of the products would be approximately 1.2. This means that almost all of the products of fission reactions have too many neutrons to be stable, which means they will themselves undergo β - decay.



Big Ideas Details

Fusion

<u>fusion</u>: the joining together of the nuclei of two atoms, accomplished by colliding them at high speeds.

Nuclear fusion reactions occur naturally on stars (such as the sun), and are the source of the heat and energy that stars produce.

On the sun, fusion occurs between atoms of deuterium (²H) to produce helium:

$$^{2}H+^{2}_{1}H\rightarrow^{4}_{2}He$$

Thermonuclear reactor



The major challenge in building nuclear fusion reactors is the high temperatures produced on the order of 10⁶–10⁹ °C. In a tokamak fusion reactor, the starting materials are heated until they become plasma—a sea of highly charged ions and electrons. The highly charged plasma is kept away from the sides by powerful electromagnets.

At the left is a schematic of the ITER tokamak reactor currently under construction in southern France.

MIT has a smaller tokamak reactor at its Plasma Science & Fusion Center. The MIT reactor is able to conduct fusion reactions lasting for only a few seconds; if the reaction continued beyond this point, the current in the electromagnets that is necessary to generate the high magnetic fields required to confine the reaction would become hot enough to melt the copper wire and fuse the coils of the electromagnet together.

After each "burst" (short fusion reaction), the electromagnets in the MIT reactor need to be cooled in a liquid nitrogen bath (-196 °C) for fifteen minutes before the reactor is ready for the next burst.

Practical Uses for Nuclear Radiation

Unit: Nuclear Chemistry

MA Curriculum Frameworks (2016): N/A (HS-PS1-8 in physics frameworks) **Mastery Objective(s):** (Students will be able to...)

• Identify & describe practical (peaceful) uses for nuclear radiation.

Success Criteria:

• Descriptions give examples and explain how radiation is essential to the particular use.

Tier 2 Vocabulary: radiation

Language Objectives:

• Explain how radiation makes certain scientific procedures possible.

Notes:

Details

While most people think of the dangers and destructive power of nuclear radiation, there are a lot of other uses of radioactive materials:

- **Power Plants**: nuclear reactors can generate electricity in a manner that does not produce CO_2 and other greenhouse gases.
- **Cancer Therapy**: nuclear radiation can be focused in order to kill cancer cells in patients with certain forms of cancer. Radioprotective drugs are now available that can help shield non-cancerous cells from the high-energy gamma rays.
- **Radioactive Tracers**: chemicals made with radioactive isotopes can be easily detected in complex mixtures or even in humans. This enables doctors to give a patient a chemical with a small amount of radioactive material and track the progress of the material through the body and determine where it ends up. It also enables biologists to grow bacteria with radioactive isotopes and follow where those isotopes end up in subsequent experiments.

Practical Uses for Nuclear Radiation

Big Ideas	Details Unit: Nuclear	Chemistry
	Irradiation of Food: food can be exposed to high-energy gamma rays in or germs. These gamma rays kill all of the bacteria in the food, but do no the food itself radioactive. (Gamma rays cannot build up inside a subs This provides a way to create food that will not spoil for months on a s store. There is a lot of irrational fear of irradiated food in the United S irradiation is commonly used in Europe. For example, irradiated milk v for months on a shelf at room temperature without spoiling.	rder to kill t make tance.) helf in a tates, but will keep
	Carbon Dating : Because ¹⁴ C is a long-lived isotope (with a half-life of 5 700 the amount of ¹⁴ C in archeological samples can give an accurate estimatheir age. One famous use of carbon dating was its use to prove that to f Turin (the supposed burial shroud of Jesus Christ) was actually made 1260 c.E. and 1390 c.E.) years), ate of :he Shroud e between
	Smoke Detectors : In a smoke detector, ²⁴¹ Am emits positively-charged algorithms particles, which are directed towards a metal plate. This steady flow of charges completes an electrical circuit. If there is a fire, smoke particle neutralize positive charges. This makes the flow of charges through the electrical circuit stop, which is used to trigger the alarm.	oha of positive es ie

Introduction: Electronic Structure

Unit: Electronic Structure

Details

Topics covered in this chapter:

The Electron	198
The Bohr Model of the Hydrogen Atom	199
The Quantum-Mechanical Model of the Atom	203
Waves	207
Electron Energy Transitions	210
Orbitals	212
Electron Configurations	217
Exceptions to the Aufbau Principle	226
Valence Electrons	228

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

HS-PS1-1 Use the periodic table as a model to predict the relative properties of main group elements, including ionization energy and relative sizes of atoms and ions, based on the patterns of electrons in the outermost energy level of each element. Use the patterns of valence electron configurations, core charge, and Coulomb's law to explain and predict general trends in ionization energies, relative sizes of atoms and ions, and reactivity of pure elements.

Big Ideas	Details	Unit: Electronic Structure						
	The Electron							
	Unit: Ele	ctronic Structure						
	 MA Curriculum Frameworks (2016): HS-PS1-1 Mastery Objective(s): (Students will be able to) Describe & explain the particle vs. wave nature of electrons. Success Criteria: 							
	 Descriptions successfully communicate accurate information about electrand their behavior. 							
	Tier 2 V	ocabulary: charge						
	Languag	e Objectives:						
	• Exp	lain scientific information about electrons.						
	Notes:							
	electron:	a small subatomic particle found outside the nucleus of an atom.						
	<u>mass</u>	= 9.11×10^{-31} kg = 9.11×10^{-28} g = $1/_{1836}$ of the mass of a proton						
	<u>charge</u>	= −1.6022 × 10 ⁻¹⁹ coulomb = −1 elementary charge						
	<u>radius</u>	= 2.8179 × 10 ⁻¹⁵ m						
	<u>electric c</u>	urrent (electricity): electrons moving from one place to another.						
	Protons a decay), b	and neutrons remain in the nucleus of their atom (except for nuclear ut electrons can be removed from one atom and added to another.						
	ion: an atom (or group of atoms that functions like a single atom) that has a electric charge because it has either gained or lost electrons.							
	Because an electron has mass (though it's very small—about $^{1}/_{1836}$ of the mass of a proton or neutron), this means electrons are particles, and all of the equations that apply to motion of solid particles also apply to electrons.							
	However, an electromagnetic wave is a wave of electricity, and electricity is made of electrons that are moving. This means that moving electrons are also waves—they move through empty space, carrying energy with them. Therefore, all of the equations that apply to waves also apply to electrons.							
	This t	means that an electron must be <u>both</u> a wave <u>and</u> a particle at the same ime.						

Unit: Electronic Structure

Details

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Describe developments that led to the Bohr model of the atom.
- Describe & explain the Bohr model of the atom.
- Explain how the quantum mechanical model of the atom grew out of the Bohr model.

Success Criteria:

- Descriptions successfully communicate developments prior to the Bohr model that were incorporated into the model.
- Descriptions successfully communicate accurate information about the Bohr model and how it describes the behavior of atoms.

Tier 2 Vocabulary: model

Language Objectives:

• Explain scientific information about the Bohr mechanical model of the atom.

Notes:

Significant Developments Prior to 1913

Atomic Theory

Significant developments in atomic theory are described in the "History of Atomic Theory" section, which begins on page 157. The most significant advances were the discovery of the electron and the planetary model of the atom.

Early Quantum Theory

<u>"Old" Quantum Theory (ca. 1900)</u>: sub-atomic particles obey the laws of classical mechanics, but that only certain "allowed" states are possible.

Big Ideas	Details		Unit: Electronic Structure
	Spectroscopy		
	Balmer Formula (1885): Swiss mat an empirical equation to relate hydrogen atom.	hematician and physici the emission lines in t	st Johann Balmer devised he visible spectrum for the
	<u>Rydberg Formula</u> (1888): Swedish generalized formula that could lines in hydrogen (and similar e	physicist Johannes Ryd describe the wave nur elements).	berg developed a nbers of all of the spectral
	There are several series of spec at different wavelengths. Rydk of integers (n ₁ and n ₂ , where n constant (now called the Rydbe	tral lines for hydrogen berg described the Baln $(1 < n_2)$, and devised a si erg constant) that relat $\frac{1}{\lambda_{HR}} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$, each of which converge ner series in terms of a pair ngle formula with a single es them.
	The value of Rydberg's constan	it is $\frac{m_e e^4}{8\varepsilon_o^2 h^3 c} = 1097373$	$31.6 \text{ m}^{-1} \approx 1.1 \times 10^7 \text{m}^{-1}$
	where m_e is the rest mass of the electrical permittivity of free splight in a vacuum.	e electron, <i>e</i> is the eler bace, <i>h</i> is Planck's const	mentary charge, \mathcal{E}_{o} is the tant, and c is the speed of

The Bohr Model of the Hydrogen Atom

Details Unit: Electronic Stru Rydberg's equation was later found to be consistent with other series discovered later, including the Lyman series (in the ultraviolet region; first discovered in 1906) and the Paschen series (in the infrared region; first discovered in 1908).

Those series and their converging wavelengths are:

Series	Wavelength	n1	n₂	
Lyman	91 nm	1	$2 \rightarrow \infty$	
Balmer	365 nm	2	$3 \rightarrow \infty$	
Pasch7en	820 nm	3	$4 \rightarrow \infty$	



Use this space for summary and/or additional notes:

Big Ideas

Details

Bohr's Model of the Atom (1913)

In 1913, Danish physicist Niels Bohr combined atomic, spectroscopy, and quantum theories into a single theory. Bohr hypothesized that electrons moved around the nucleus as in Rutherford's model, but that these electrons had only certain allowed quantum values of energy, which could be described by a quantum number (n). The value of that quantum number was the same n as in Rydberg's equation, and that using quantum numbers in Rydberg's equation could predict the wavelengths of light emitted when the electrons gained or lost energy by moved from one quantum level to another.



Bohr's model gained wide acceptance, because it related several prominent theories of the time. The theory worked well for hydrogen, giving a theoretical basis for Rydberg's equation. Bohr defined the energy released when an electron descended to an energy level using an integer quantum number (n) and Rydberg's constant:

$$E_n = -\frac{R_H}{n^2}$$

Bohr received the Nobel Prize in physics in 1922 for his contributions to quantum and atomic theory.

Although the Bohr model worked well for hydrogen, the equations could not be solved exactly for atoms with more than one electron, because of the additional

effects that electrons exert on each other (e.g., via the Coulomb force, $F_e = \frac{kq_1q_2}{r^2}$).

The Quantum-Mechanical Model of the Atom

Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Describe & explain the quantum mechanical model of the atom.

Success Criteria:

• Descriptions successfully communicate accurate information about the quantum mechanical model of the atom and how it describes the behavior of atoms.

Tier 2 Vocabulary: mechanical

Language Objectives:

• Explain scientific information about the quantum mechanical model of the atom.

Notes:

<u>quantum</u>: a discrete increment (plural: *quanta*) If a quantity is *quantized*, it means that only certain values for that quantity are possible.

Because an electron behaves like a wave, it can only absorb energy in quanta that correspond to exact multiples of its wavelength.

Neils Bohr was the first to realize that because atomic spectral emissions are quantized, electron energy levels must also be quantized. (See the "Historical Development of Atomic Theory" section on page 158 for more details.)

Recall that in the Bohr model of the atom, an electron's (quantum) energy level determined its distance from the nucleus.



The Quantum Mechanical Medel of the Atom

	The Quantum-Wechan	lical wodel of the	Atom Page: 20
Big Ideas	Details		Unit: Electronic Structur
	What actually happens is not that level, the Bohr model may predict electron is also a particle, so it has In 1925, Austrian physicist Erwin S a unique wave function, the energy mathematical solutions to a wave equation to construct a probabilit atom is the basis for the modern	simple. If an electron is t its <i>average</i> distance fror s some freedom to move Schrödinger found that by gies of the electrons could e equation. The use of Sch and for where the elect quantum-mechanical mo	in a particular energy n the nucleus, but the closer or farther. y treating each electron a d be predicted by the hrödinger's wave ctrons can be found in an del of the atom.
	To understand the probability map, it is important to realize that because the electron acts as a wave, it is detectable when the amplitude of the wave is nonzero, but not detectable when the amplitude is zero. This makes it appear as if the electron is teleporting from place to place around the atom. If you were somehow able to take a time-lapse picture of an electron as it moves around the nucleus, the picture might look something like the diagram to the right, where each dot is the location of the electron at an instant in time.	A Probability Density of Electron	tomit ucleus
	Notice that there is a region close found, and a ring a little farther of finding the electron. As you get fa equation predicts different shape of high probability are called "orb originally suggested by the planet	to the nucleus where the ut where there is a much arther and farther from tl s for these probability dis itals," because of their re ary model.	e electron is unlikely to be higher probability of he nucleus, Schrödinger's stributions. These regions elation to the orbits
	The quantum mechanical model of quantized energy levels and proba	of the atom is based on th abilities.	is combination of

The Quantum-Mechanical Model of the Atom Page: 206

Big Ideas	Details Unit: E	ectronic Structure
	The claims of the quantum mechanical model of the atom are:	
	 The electrons orbiting an atom behave like waves as well as their behavior can be described by Schrödinger's wave equa integer solutions called quantum numbers. 	particles, and tion, which has
	 The energies and therefore locations of electrons within an a determined exactly, but there are regions with a high probal electron (called "orbitals"), and other regions with a low pro an electron. 	atom cannot be bility of finding an bability of finding
	 The energy of each electron (and therefore its probable local described by a unique set of quantum numbers. No two electron is un the same energies, which are described by the electron is un quantum numbers, which means no two electrons can have set of quantum numbers. This is called the <i>Pauli exclusion pu</i> after the Swiss-American physicist Wolfgang Pauli. 	tion) can be ctrons can have ique set of the exact same <i>rinciple,</i> named
	Electrons move within their orbitals at speeds near the spee	d of light.
	 An electron is constrained to stay within its orbital because of the electron absorbs energy, it can move to a higher-energy unoccupied lower-energy orbital is available, the electron can (in the form of a photon, which can be observed as light) and lower-energy orbital. 	of its energy. If orbital. If an n release energy d move to the
	 The shape of any given orbital (the region where there is hig finding an electron) depends on all of the forces that affect t Some of these forces relate to the energy characteristics of t orbital, but other forces can include electrostatic repulsion of within the atom, electrostatic repulsion of the electrons in ic atoms, or the sharing of electrons in a covalently-bonded ato the shapes of orbitals are continuously changing as all of the the atom repel one another as they move at near-light speed 	h probability of hat electron. he specific f other electrons onically-bonded om. In real atoms, electrons within d.
	Use this space for summary and/or additional notes:	

_

vvaves	Page: 207				
Big Ideas Details Unit: Ele	ctronic Structure				
Waves	Waves				
Unit: Electronic Structure					
MA Curriculum Frameworks (2016): HS-PS1-1					
Mastery Objective(s): (Students will be able to)					
 Explain what waves are and how they propagate. 					
 Describe the relative energies of different waves based on the and positions within the electromagnetic spectrum. 	ir frequencies				
 Calculate the frequency and wavelength of electromagnetic v 	aves.				
Success Criteria:					
 Descriptions are accurate and backed up by evidence. 					
 Calculations are correct. 					
 Algebra and rounding to appropriate number of significant fig 	ures is correct.				
Tier 2 Vocabulary: wave, spectrum					
Language Objectives:					
 Explain scientific information about waves and the electroma 	netic spectrum.				
Notes:					
wave: an energy disturbance that travels from one place to anothe	wave: an energy disturbance that travels from one place to another.				
medium: the substance that a wave travels through. Electromagne (including light) can travel without a medium.	<u>medium</u> : the substance that a wave travels through. Electromagnetic waves (including light) can travel without a medium.				
 The wave travels through the medium. 					
 All (or nearly all) of the energy passes through the medium— doesn't absorb it. 	 All (or nearly all) of the energy passes through the medium—the medium doesn't absorb it. 				
Some examples of waves:					
Type of Wave Medium					
sound air (or water, sol	ds)				
ocean water					
electromagnetic (<i>e.g.,</i> light, radio) none					

Big Ideas	Details Unit: Electronic Structure
	<u>wavelength</u> (λ): the length of the wave, measured from a specific point in the wave to the same point in the next wave. unit = distance (m, cm, nm, <i>etc.</i>)
	<u>frequency</u> : (f or v) the number of waves that travel past a point in a given time. Symbol = f; unit = $\frac{1}{time}$ = Hz
	speed = λf
	Electromagnetic waves (such as light, radio waves, etc.) travel at a constant speed—the speed of light. The speed of light is a constant, and is denoted by the letter " c " in equations.
	$c = 3.00 \times 10^8 \text{ m/s} = 186,000 \text{ miles per second}$
	The energy (E) that a wave carries equals a constant times the frequency. (Think of it as the number of bursts of energy that travel through the wave every second.) For electromagnetic waves (including light), the constant is Planck's constant (named after the physicist Max Planck), which is denoted by a script h in equations. So the equation is:
	E = hf
	where $h = 6.63 \times 10^{-34}$ J·s = Planck's constant
	Louis de Broglie: French physicist. Showed that any object with momentum (<i>i.e.,</i> has mass and is moving) creates a wave as it moves.
	Large objects with a lot of momentum (such as people) create waves with wavelengths that are far too small to detect.
	Small objects (such as electrons) create waves with wavelengths in the visible part of the spectrum. This is why we can see the light produced by electrons as they move.



Use this space for summary and/or additional notes:

Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Explain the meaning of the lines in emission spectra.

Success Criteria:

• Descriptions include colors, approximate wavelengths, and relationship between number of lines and number of electrons.

Tier 2 Vocabulary: emission, spectrum, ground, excited

Language Objectives:

• Explain scientific information about energy transmissions and line spectra.

Notes:

Details

<u>quantum</u>: a discrete quantity of energy that cannot be divided.

ground state: the lowest available energy level for an electron.

excited state: a higher energy level than the ground state.

<u>emission spectrum</u>: the wavelengths (colors) of light emitted by an element when its electrons are "excited" (raised to a higher energy state) and then allowed to return to the ground state.

spectroscope: a device that separates colors of light based on their wavelengths.

Electron Energy Transitions

DetailsUnit: Electronic StructureIf you were to look at a glass tube filled with hydrogen gas that was energized with
electricity, the gas in the tube would appear to be blue, because the electrons are
energized, and the energy of the light they emit as they return to the n = 2 quantum
energy level corresponds with a blue color.If you were to split the light emitted by hydrogen into its component colors using a

If you were to split the light emitted by hydrogen into its component colors using a spectroscope, you would see the following:



Atoms with more electrons have a larger number of possible transitions, each with different energies. This results in more lines in their emission spectrum, as with iron:



photon: a single "piece" (particle/wave) of light.

<u>luminescence</u>: light that is not generated by high temperatures alone. (In fact, it usually occurs at low temperatures. Causes include electrical energy and chemical reactions.)

- <u>fluorescence</u>: a type of luminescence that occurs when electrons of an element are excited and return immediately to the ground state, giving off a photon. The wavelength of the photon given off is usually different from the wavelength of the photon used to raise the electrons to the excited state.
- <u>phosphorescence</u>: "glow-in-the-dark" luminescence—a type of luminescence that occurs when electrons are excited, but cannot return directly to the ground state. The indirect path is slower, which cause the material to "glow" for a longer period of time (in some cases, hours).
- triboluminescence: a form of luminescence in which light is generated by breaking asymmetrical bonds in a crystal. In the case of wintergreen Life Savers, when the sugar crystals are crushed, the positive and negative charges get separated. The voltage between them causes a spark. Normally, these sparks are in the ultraviolet part of the spectrum and cannot be seen; however, wintergreen oil (methyl salicylate) is fluorescent. It absorbs the photons of ultraviolet light and emits photons of blue light, which we can see.

Use this space for summary and/or additional notes:

Big Ideas

Orbitals

Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Explain the energy hierarchy of quantum levels, sub-levels, and orbitals.
- Explain how the hierarchy of quantum levels, sub-levels and orbitals corresponds with positions on the periodic table of the elements.

Success Criteria:

• Descriptions relate principal quantum number to period and sub-level to region of the periodic table.

Tier 2 Vocabulary: level, sub-level

Language Objectives:

• Explain the energy hierarchy of quantum levels, sub-levels and orbitals.

Notes:

Details

Big Ideas

<u>orbital</u>: a region in an atom (outside the nucleus) with a high probability of finding an electron.

These regions are called *orbitals* because these regions are what replaced the spherical orbits in the much simpler planetary model.

Note that orbitals are not physical objects with boundaries. They are simply the space that an electron moves around in because its energy and the external forces on it. When an electron approaches the "boundary" of its orbital, the forces pulling it back are strong enough to overcome the energy that the electron has, and it cannot get farther away.

The locations and geometric shapes of these orbitals are the solutions to the wave equation, a complex mathematical formula that would require mathematics far too advanced for a high school class. Instead, we will categorize orbitals using a hierarchy that is divided according to energy levels.

	Orbitals	Page: 213			
Big Ideas	Details	Unit: Electronic Structure			
	Energy Level Hierarchy				
	Electrons have energy levels, which roughly correspond Rutherford-Bohr model. Each of those levels has one o Each sublevel has one or more orbitals, and each of the two electrons. It's easiest to think of the hierarchy as a	d to the values of <i>n</i> in the r more kinds of sublevels. ose orbitals can hold up to in outline:			
	 energy level (1, 2, 3,) a. sub-level (s, p, d, f,) i. orbital a. individual electron 				
	Energy Level				
	The main or principal level is a measure of total distance levels are numbered 1-7. The periodic table of the eler Period (row number) that an element is in equals the e that has the most energy.	e from the nucleus. The nents is arranged so that the nergy level of the electron			
	For example, helium is in period (row) #1. That means which means both of its electrons have to be in level 1.	its highest energy level is 1,			
	Sulfur is in period (row) #3, which means it has electror	ns in levels 1, 2, and 3.			

Sub-Levels & Or

Details

Big Ideas

There are four types of sub-levels. Each type has a specific number of orbitals with specific shapes. Each of those orbitals can hold up to 2 electrons.

There are four kinds of sub-levels: **s**, **p**, **d**, and \mathbf{f}^* . The shapes of their orbitals are:

type of sub-level	shape(s) of orbital(s)	total # of orbitals	total # of electrons
S		1	2
р		3	6
d		5	10
f		7	14
Notice that e	ach sub-level has an odd number of orbitals.		

^{*} The letters come from words that described the characteristics of the atomic spectra. **s** stood for "sharp," **p** for "principal," **d** for "diffuse," and **f** for "fundamental."

Arrangement of Levels and Sublevels

Each energy level in an atom has at least one of these kinds of sub-levels—each level contains an *s* sub-level, each level starting with level 2 contains *s* and p sub-levels, each level starting with level 3 contains *s*, *p*, and d sub-levels, *etc*.

Level	1	2	3	4	5	6	7
# of sub-levels	1	2	3	4	5	6	7
types of sub-levels	S	s, p	s, p, d	s, p, d, f	s, p, d, f, (g)	s, p, d, f, (g), (h)	s, p, d, f, (g), (h), (i)

Note that g, h, and i are in parentheses because no atom is large enough that it actually has any electrons in g, h, or i sublevels, but mathematically we know that those sub-levels will exist if we "discover" (create) a large enough atom.

Note also that these sub-levels overlap. For example, the levels and sublevels in a sodium (Na) atom might look like the following:



Use this space for summary and/or additional notes:

Big Ideas

Details

Big Ideas Details Sub-levels and the Periodic Table Note that the sub-level of the electron that has the highest energy corresponds with the location of the element on the periodic table: s 15 1s 25 2р Зр 3s **4**5 3d **4**p p h 5s 4d 5p ÷ 6s 5d 6р 7s 6d 7p



Notice, for example, that the "s section" of the periodic table is two columns wide. This is because each **s** sub-level has one orbital that can hold two electrons.

type of sub-level	# orbitals	×	2 = # electrons		# co O	olumns n P.T.	
S	1			2	•	-	2
р	3			6	•		6
d	5			10	←		10
f	7			14	←		14
Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Write the ground-state electron configuration for any element on the periodic table.

Success Criteria:

- Levels and sublevels are in the correct order.
- Each sublevel has the correct number of electrons.

Tier 2 Vocabulary: level, spin

Language Objectives:

• Explain the parts of an electron configuration.

Notes:

The electron configuration for an element is a list of all of the energy sub-levels that have electrons in them (in the ground state) in order from lowest to highest energy, and the number of electrons they contain.

For example, consider a neutral nitrogen atom with its seven electrons.

- The first two electrons occupy the 1s sublevel—the one with the lowest energy. We denote these two electrons as 1s².
- The next two electrons occupy the 2s sublevel. We denote these two electrons as 2s².
- The last three electrons are in the 2p sublevel. We denote these three electrons as $2p^3$.
- The complete electron configuration for nitrogen is therefore 1s² 2s² 2p³.

If this already makes sense, great! The next few pages explain where these numbers come from in more detail.





^{*} For the purpose of electron configurations, helium should be in the "s" block, next to hydrogen.

Use this space for summary and/or additional notes:

Big Ideas

Big Ideas	Details Unit: Electronic Structure
	Writing Electron Configurations
	An element has electrons that correspond with <u>each</u> of the available slots, from the beginning of the periodic table (where hydrogen is located) up to where that element is located.
	If we were to represent an electron as an arrow, we could represent two electrons
	in a 1s sub-level like this: $\frac{\uparrow\downarrow}{1s}$. The 1s sub-level has one orbital, which is
	represented by the one blank. The two electrons are represented as arrows. Because two electrons sharing an orbital have opposite spins, we represent them with one arrow pointing up and the other arrow pointing down.
	We could represent five electrons in a 2p orbital like this: $\frac{\uparrow\downarrow}{2p} \stackrel{\uparrow\downarrow}{-}$. The 2p sub-
	level has 3 orbitals, represented by the 3 blanks. Two of those orbitals have two electrons in them, and the third one has only one electron.
	We could represent all 13 of the electrons in aluminum like this:
	$\uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow$
	$\frac{1}{1s} \frac{2}{2s} \frac{2}{2p} \frac{3}{3s} \frac{3}{3p}$
	This diagram shows the <u>electron configuration</u> of aluminum.
	electron configuration: a description of which levels and sub-levels the electrons in an element are occupying.
	Notice that we have to show all three of the orbitals (blanks) in the 3p sub-level, even if some of those orbitals don't have any electrons in them.
	ground state: when all of the electrons in an atom are in the lowest-energy sublevel that has an available "slot".
	Pauli Exclusion Principle: every electron in an atom has a different quantum state from every other electron. In plain English, this means that something has to be different about each electron, whether it's the level, sub-level, which orbital it's in, or its spin.
	aufbau principle: in the ground state, each electron in an atom will occupy the lowest available energy state. This means that you start with the lowest sub-level (1s) and work your way up until you've placed all the electrons.

g Ideas	Details Unit: Electronic Structu
	Hund's Rule: electrons don't pair up in orbital until they have to. (This is kind of li
	siblings not wanting to share a room if there's an empty room available.) For
	example, the electron configuration for nitrogen would be:
	$\uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow$
	Wrong: $\frac{1}{1}$ $\frac{1}{2}$ $\frac{1}{2}$
	1s 2s 2p
	$\uparrow \downarrow \uparrow \downarrow \uparrow \uparrow \uparrow \uparrow \bullet \bullet$
	Right: $\frac{1}{1c}$ $\frac{1}{2c}$ $-\frac{1}{2n}$
	13 23 2p
	If you don't need to draw every electron, you can use a shorter form, in which you
	just write the level and sub-level, and use a superscript for the number of electror
	in the sub-level.
	$\uparrow \downarrow$
	For example, — would become 15° , and — — — would become $2p^{\circ}$.
	The electron configuration for eluminum would go from the orbital notation
	version:
	$\uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow$
	$\frac{1}{1s}$ $\frac{1}{2s}$ $\frac{1}{2n}$ $\frac{1}{3s}$ $\frac{1}{3n}$
	10 20 2p 00 0p
	to the "atom doud" wavelend
	to the standard version:
	$1s^2 2s^2 2n^6 3s^2 3n^1$
	13 23 2h 33 3h

Details						U					Un	it: El	lectr	onic	Stru	cture	9
The shorte	r versio	n can	still	get t	edio	usly	long	g for	elem	ents	s wi	th a	lot c	of ele	ctro	ns.	-
For examp	le, the e	lectro	on co	onfig	urati	on f	or go	old (A	Au) is	:							
	1s ²	2s ² 2	2p ⁶ 3s	s² 3p	⁶ 4s ²	3d ¹⁰	⁰ 4p ⁶	5s ² 4	1d ¹⁰ !	5p ⁶ 6	5s ² 4	4f ¹⁴ 5	5d ⁹				
To shorten	this eve	en mo	ore, y	/ou'r	e all	owe	d to	use t	he e	leme	ent	in th	ne la	st co	lumi	n of	
a row as an	n apprev	/latio	n tor	ali o	T the	eie	ctror	ns th	roug	n the	e er	10 01	r tha	t rov	V.		
In our exar	nple, go	ld (A	u) is i	n th	e 6tł	n row	voft	the p	erio	dic ta	able	9:					
L dd																	18 VIII A
1 1s 2 1 1 A	·										_	13 III A	14 IV A	15 V A	16 VI A	17 VII A	
2 25											2p						
3 3s	3 III B	4 IV B	5 V B	6 VI B	7 VII B	8 VIII B	9 VIII B	10 VIII B	11 I B	12 II B	3p						
4 4s	3d										4p						
5 55	4d										5p						Xe
^{6 6s} 6s ²	⁵₫ 5d ¹	5d ²	5d ³	5d ⁴	5d⁵	5d ⁶	5d ⁷	5d ⁸	Au		6р						
7 7s	6d										7р						
lanthani (rare earth me	des 4f						4	f ¹⁴									
actini	des 5f																
This means			ed to	, star	t fro	m ve					nd	of th			 (5	th	
row, and a	dd on o	nly th	ie pa	rts th	nat c	ome	afte	er Xe	. Thi	s giv	es i	us:	ie pi	evior	כ) כנ	1	
	$(1s^2)$	$2s^2 2$	p ⁶ 3s	² 3p ⁶	4s ²	3d ¹⁰	4p ⁶	5s ² 4	d ¹⁰ 5	0 [€] 0	5s ² 4	4f ¹⁴ !	5 d ⁹				
							T				-						
							↓										
No	ble gas	confi	gurat	tion:		[]	Xe] 6	5s ² 4f	¹⁴ 5c	1 ⁹							
This notation is called						onfi	gura	<u>tion</u> ,	beca	ause	the	e ele	mer	its in	the	last	
column (the ones you start fro							ed th	ne no	ble g	gases	5.						
	Details The shorte For examp To shorten a row as ar In our exar $\frac{8}{2}$ $\frac{1}{1A}$ $\frac{2}{1A}$ 1 15 $\frac{2}{1A}$ $\frac{2}{1A}$ 2 25 $\frac{1}{1A}$ $\frac{2}{1A}$ 3 35 $\frac{1}{1A}$ $\frac{2}{1A}$ 4 46 $\frac{1}{1A}$ $\frac{2}{1A}$ 5 55 $\frac{1}{1A}$ $\frac{2}{1A}$ 7 75 $\frac{1}{1A}$ $\frac{2}{1A}$ 6 65 $\frac{6s^2}{7}$ $\frac{1}{7}$ $\frac{1}{75}$ $\frac{1}{1A}$ $\frac{1}{1A}$ $\frac{2}{1A}$ $\frac{1}{1A}$ $\frac{2}{1A}$ $\frac{1}{1A}$ $\frac{2}{1A}$ $\frac{2}{1A}$ $\frac{1}{1A}$ $\frac{2}{1A}$	Details The shorter version For example, the end 1s ² To shorten this even a row as an abbreve In our example, go $\frac{1}{2}$ $\frac{1}{1}$ $\frac{1}{1}$ $\frac{1}{1}$ $\frac{1}{2}$ $\frac{1}{2}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{5}$ $\frac{1}{5}$ $\frac{1}{6}$ $\frac{1}{7}$ $\frac{1}{7}$ $\frac{1}{7}$ This means we're a row, and add on o $1s^2$ Noble gas This notation is call column (the ones y	Details The shorter version can For example, the electron $1s^2 2s^2 2$ To shorten this even model a row as an abbreviation In our example, gold (A $\frac{9}{6}$ $\frac{1}{14}$ $1 \frac{1}{15}$ $\frac{2}{14}$ $3 \frac{3}{14}$ $\frac{3}{14}$ 3	Details The shorter version can still. For example, the electron co $1s^2 2s^2 2p^6 3s^2$ To shorten this even more, y a row as an abbreviation for In our example, gold (Au) is is $\frac{9}{6}$ $\frac{1}{14}$ $1 \frac{1}{14}$ $\frac{1}{14}$ $\frac{1}{14$	Details The shorter version can still get t For example, the electron config $1s^2 2s^2 2p^6 3s^2 3p^6$ To shorten this even more, you'r a row as an abbreviation for all o In our example, gold (Au) is in the $\frac{36}{2}$ $\frac{1}{4}$ $\frac{1}{2} 2s^2$ $\frac{1}{4}$ $$	Details The shorter version can still get tedio For example, the electron configurati $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ To shorten this even more, you're allow a row as an abbreviation for all of the In our example, gold (Au) is in the 6th $\frac{5}{2}$ $\frac{1}{4$	Details The shorter version can still get tediously For example, the electron configuration for $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$ To shorten this even more, you're allowed a row as an abbreviation for all of the elect In our example, gold (Au) is in the 6th row $\frac{3}{2}$ $\frac{1}{4}$	Details The shorter version can still get tediously long For example, the electron configuration for get $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ To shorten this even more, you're allowed to a row as an abbreviation for all of the electron In our example, gold (Au) is in the 6th row of f $\frac{3}{9}$ $\frac{1}{10}$ $\frac{1}{9}$ 1	Details The shorter version can still get tediously long for For example, the electron configuration for gold (<i>A</i> 1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ¹⁰ 4p ⁶ 5s ² <i>A</i> To shorten this even more, you're allowed to use t a row as an abbreviation for all of the electrons th In our example, gold (Au) is in the 6th row of the p	Details The shorter version can still get tediously long for elem For example, the electron configuration for gold (Au) is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 5d^2 5d^2 5d^2 5d^2 5d^2 5d^2 5d^2 5d^2$	Details The shorter version can still get tediously long for elements For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6$ To shorten this even more, you're allowed to use the element a row as an abbreviation for all of the electrons through the In our example, gold (Au) is in the 6th row of the periodic to $\frac{9}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ This means we're allowed to start from xenon (Xe) at the errow, and add on only the parts that come after Xe. This giv $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6$ Noble gas configuration: [Xe] $6s^2 4f^{14} 5d^9$ This notation is called the noble gas configuration, because column (the ones you start from) are called the noble gases	Details Un The shorter version can still get tediously long for elements wi For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4d^{1$	Details Unit: E The shorter version can still get tediously long for elements with a For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 9$ To shorten this even more, you're allowed to use the element in th a row as an abbreviation for all of the electrons through the end of In our example, gold (Au) is in the 6th row of the periodic table: $\frac{9}{2}$ $\frac{1}{4}$	Details Unit: Electric The shorter version can still get tediously long for elements with a lot of For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^9$ To shorten this even more, you're allowed to use the element in the <i>la</i> a row as an abbreviation for all of the electrons through the end of tha In our example, gold (Au) is in the 6th row of the periodic table: $\frac{3}{2}$ $\frac{1}{4}$ 	Details Unit: Electronic The shorter version can still get tediously long for elements with a lot of ele For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^9$ To shorten this even more, you're allowed to use the element in the <i>last</i> co a row as an abbreviation for all of the electrons through the end of that row In our example, gold (Au) is in the 6th row of the periodic table: $\frac{3}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ $\frac{1}{4}$ This means we're allowed to start from xenon (Xe) at the end of the previou row, and add on only the parts that come after Xe. This gives us: $(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6) 6s^2 4f^{14} 5d^9$ This notation is called the <u>noble gas configuration</u> , because the elements in column (the ones you start from) are called the noble gases.	Details Unit: Electronic Strue The shorter version can still get tediously long for elements with a lot of electro For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^9$ To shorten this even more, you're allowed to use the element in the <i>last</i> column a row as an abbreviation for all of the electrons through the end of that row. In our example, gold (Au) is in the 6th row of the periodic table: $\frac{3}{2}$ $\frac{1}$	Details Unit: Electronic structure The shorter version can still get tediously long for elements with a lot of electrons. For example, the electron configuration for gold (Au) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^9$ To shorten this even more, you're allowed to use the element in the <i>last</i> column of a row as an abbreviation for all of the electrons through the end of that row. In our example, gold (Au) is in the 6th row of the periodic table: $\frac{9}{4}$ $\frac{1}$

	Electron Configurations	Page: 223
Big Ideas	Details Unit: Electronic	Structure
	Homework Problems	
	Give the electron configuration (orbital notation—with the arrows) for eac following elements:	h of the
	1. carbon	
	2. potassium	
	3. silicon	
	4. silver	
	For each of the following electron configurations, name the element.	
	5. $\frac{\uparrow\downarrow}{1s} \frac{\uparrow\downarrow}{2s} \frac{\uparrow\downarrow}{2p} \frac{\uparrow\downarrow}{2p}$	
	$6. \frac{\uparrow\downarrow}{1s} \frac{\uparrow\downarrow}{2s} \frac{\uparrow\downarrow}{2p} \frac{\uparrow\downarrow}{3s} \frac{\uparrow\downarrow}{3p} \frac{\uparrow\downarrow}{3p} \frac{\uparrow\downarrow}{4s} \frac{\uparrow\downarrow}{4s} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{4s} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{4s} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{1} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{1} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{1} \frac{\uparrow\downarrow}{3d} \frac{\uparrow\downarrow}{1} \frac{\downarrow\downarrow}{1} \frac{\downarrow\downarrow}{1} $	
	Each of the following electron configurations has something wrong with it. For each one: • State what the mistake is. • Re-write the electron configuration correctly, keeping the total number electrons the same. 7. $\frac{\uparrow\downarrow}{1s} \frac{\uparrow\downarrow\uparrow\uparrow}{2p} \frac{\uparrow}{2p}$	per of

Big Ideas	Details Unit: Electronic Stru
	8. $\frac{\uparrow\downarrow}{1s} \frac{\uparrow\downarrow}{2s} \frac{\uparrow\downarrow}{2p}$
	9. $\frac{\uparrow\downarrow}{1s} \frac{\uparrow\downarrow}{2s} \frac{\uparrow\downarrow}{2p} \frac{\uparrow\downarrow}{3s} \frac{\uparrow\downarrow}{3p} \frac{\uparrow\downarrow}{3p}$
	For each of the following elements, give the "standard" electron configuration (<i>e.g.</i> , 1s ² 2s ² 2p ⁶ 3s ¹). 10. boron (B)
	11. phosphorus (P)
	12. vanadium (V)
	13. strontium (Sr)
	For each of the following electron configurations, give the element. 14. 1s ² 2s ² 2p ⁶ 3s ²
	15. 1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ¹⁰ 4p ⁶ 5s ² 4d ⁶

Big Ideas	Details Unit: Electronic Structure	е
	For each of the following elements, give the "noble gas" electron configuration (e.g., [Ar] $4s^2 3d^5$).	
	16. zirconium (Zr)	
	17. platinum (Pt)	
	18. dysprosium (Dy)	
	19. gallium (Ga)	
		-

Exceptions to the Aufbau Principle

Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Predict which elements are likely to be exceptions to the aufbau principle.
- Explain why exceptions to the aufbau principle occur as you approach the middle and end of the d and f sub-levels.

Success Criteria:

• Predictions match observed electron configurations.

Language Objectives:

• Explain exceptions to the aufbau principle.

Notes:

Details

Remember from Hund's Rule that electrons like to spread out.

Atoms are the most stable when their electrons are the most evenly distributed within the atom's energy levels and sub-levels. This means that elements with completely filled principal (numbered) energy levels are the most stable.

- The "noble gases" (the last column of the periodic table) already have all of their principal energy levels completely filled with electrons. This makes them very stable, because they do not need to react with other atoms to get their electrons into a more stable configuration. This is why noble gases almost never react with anything.
- Other elements gain, lose, or share electrons (in chemical reactions) in order to end up with electron configurations that are like the nearest noble gas on the periodic table.

Atoms with p, d, and f sub-levels that are exactly half full are more stable than atoms with slightly more or fewer electrons in their p, d, and f sub-levels. This makes those atoms slightly more stable (and therefore less reactive) than other atoms. For example:

- Nitrogen ([He] 2s² 2p³), which has an exactly half-filled 2p sub-level, is chemically less reactive than oxygen ([He] 2s² 2p⁴).
- Manganese ([Ar] 4s² 3d⁵), which has an exactly half-full 3d sub-level, is chemically less reactive than iron ([Ar] 4s² 3d⁶).

Exceptions to the Aufbau Principle Page 227

	Exceptions to the Adibad Finiciple Page: 227
Big Ideas	Details Unit: Electronic Structure
	In fact, elements with a d or f sub-level that is one electron away from being half full will usually "borrow" one electron from the nearest s sub-level, because the half-filled d or f sub-level is more stable than the full s sub-level.
	 Chromium "borrows" one of its 4s electrons to make its 3d sub-level exactly half full. This means that instead of having predicted electron configuration of [Ar] 4s² 3d⁴, it is observed to have the electron configuration [Ar] 4s¹ 3d⁵. This happens because a half-filled 4s sub-level plus a half-filled 3d sub-level is more stable than a completely filled 4s sub-level plus a 3d sub-level with 4 electrons in it.
	 Copper "borrows" one of its 4s electrons to make its 3d sub-level completely full. This means that instead of having predicted electron configuration of [Ar] 4s² 3d⁹, it is observed to have the electron configuration [Ar] 4s¹ 3d¹⁰. Again, this happens because a half-filled 4s sub-level plus a completely filled 3d sub-level is more stable than a completely filled 4s sub-level plus a 3d sub-level with 9 electrons in it.
	There are a significant number of other exceptions to the aufbau principle. Clearly, atoms do not care about the periodic table!

Unit: Electronic Structure

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Determine the number of valence electrons for representative elements.
- Draw Lewis dot diagrams for representative elements.

Success Criteria:

Details

- Elements are drawn with the correct number of valence electrons.
- Dots representing electrons are spread around the element symbol in an appropriate fashion.

Language Objectives:

• Explain what valence electrons are and how to determine how many an element has.

Notes:

<u>valence electrons</u>: the outer electrons of an atom that are available to participate in chemical reactions.

In most atoms, these are the electrons in the s and p sub-levels of the highest (numbered) energy level.

For example, phosphorus (P) has the electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^3$, or [Ne] $3s^2 3p^3$. The highest energy level is level 3.

The $3s^2 3p^3$ at the end of its electron configuration tells us that phosphorus has 2 electrons in the 3s sub-level plus 3 in the 3p sub-level, for a total of 5 electrons in level 3. This means that phosphorus has 5 valence electrons.

Note that only electrons in s and p sub-levels can be valence electrons. For example, arsenic (As) has the electron configuration [Ar] $4s^2 3d^{10} 4p^3$. The highest energy level is 4, so only the electrons in level 4 count. Arsenic has 2 electrons in the 4s sub-level, and 3 electrons in the 4p sub-level, for a total of 5 valence electrons. The 10 electrons in the 3d sub-level are <u>not</u> in the highest level, so they don't count.

Valence Electrons

Big Ideas	Details	Unit: Electronic Structure
	Recall that full sub-levels give an atom extra elements in the last column of the periodic t because all of their sub-levels are filled. This with other elements.	stability. This means noble gases (the cable) are the most stable elements s is why noble gases almost never react
	Because noble gases have all sub-levels filled shells. Helium has 2 valence electrons (beca other noble gases have 8 valence electrons level is full with 2 electrons, and their highes 6 electrons, for a total of 8.)	d, this means they have "full" valence suse it has only a 1s sub-level), and all (because their highest-numbered s sub- st-numbered p sub-level is full with
	For other elements, <i>the atoms can become to with filled valence shells</i> , which would give to as a noble gas.	<i>much more stable if they can form ions</i> he ion the same electron configuration
	For example, phosphorus ([Ne] 3s ² 3p ³) has valent shell by gaining 3 more electrons to fi the same electron configuration as argon), o give it the same electron configuration as ne 3 electrons than to lose 5, phosphorus is more means it's most likely to form an ion with a	5 valence electrons. It could have a full ill its 3p sub-level (which would give it or by losing 5 electrons (which would eon). Because it is easier to gain ost likely to gain 3 electrons, which -3 charge.
	Potassium ([Ar] 4s ¹) has only one valence ele electron (which would give it the same elect 7 electrons (which would give it the same el Because it is easier to lose 1 electron than to 1 electron, which means it's most likely to fo	ectron. Potassium could either lose 1 fron configuration as argon), or gain ectron configuration as krypton). o gain 7, potassium is most likely to lose form an ion with a +1 charge.
	Transition Metals	
	Because the energy of an <i>s</i> sub-level is so clot the next lower energy level, transition meta these <i>s</i> and <i>d</i> sub-levels. This means they ca electrons, depending on the situation. For e configuration [Ar] $4s^2 3d^9$, or [Ar] $4s^1 3d^{10}$, m or two valence electrons. This explains why a +1 ion, and other times a +2 ion.	ose to the energy of the <i>d</i> sub-level of Is can easily shift electrons between an have different numbers of valence example, copper can have the electron beaning that copper can have either one copper is observed to sometimes form
	Use this space for summary and/or addition	al notes:

Valence Electrons

Big Ideas	Details Unit: Electronic Structure							
	Group Numbers							
	You can read the number of valence electrons that an element has directly from the periodic table, using the group numbers. For the "representative elements" (s and p block elements), the number of valence electrons is the last digit of the group number. Transition metals generally have two valence electrons, though there are exceptions. (See the section on "Exceptions to the Aufbau Principle" starting on page 226 for an explanation.)							
	Lewis Dot Diagrams							
	A Lewis dot diagram is a representation of an element surrounded by its valence electrons. The diagram consists of the element symbol (from the periodic table), with dots on the top, bottom, and sides representing the <i>s</i> and <i>p</i> sub-levels of its valence shell.							
	For example, aluminum has 3 valence electrons. The orbital-notation electron configuration for aluminum is:							
	$\uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow \downarrow \ \uparrow$							
	$\frac{1}{1s} \frac{2}{2s} \frac{2}{2p} \frac{3}{3s} \frac{3}{3p}$							
	Its Lewis dot diagram is ·AI:							
	Notice that it shows three dots representing the 3 valence electrons.							
	The dots are placed in singles or pairs on the top, bottom, left, and right of the element symbol. The convention is to place the first two valence electrons (the ones in the s sub-level) to the right of the element symbol, and the remaining valence electrons (the ones in the p sub-levels) on the top, left, and bottom. Start with one dot on the top, left, and bottom, and then pair them up one at a time. (This corresponds with Hund's Law, which says that electrons in the p sub-level do not pair up until they have to.)							
	In our example, the Lewis dot diagram for aluminum has two dots on the right representing the two electrons in the 3s sub-level, and one dot on the left for the one electron in the 3p sub-level.							

Valence Flectrons

	valend	ce Electrons	Page: 231
Big Ideas	Details		Unit: Electronic Structure
	Nitrogen has 5 valence electro	ns. Its orbital-notation ele	ectron configuration is:
		$\stackrel{\uparrow\downarrow}{-} \stackrel{\uparrow\downarrow}{-} \stackrel{\uparrow}{-} \stackrel{\downarrow}{-} $	
		1s 2s 2p	
	Its Lewis dot diagram would be	e •N:	
	Again, notice that there are 2 deach on the top, bottom, and l the 2p sub-level.	dots on the right for the 2s left sides for the one elect	s sub-level, and one dot ron in each of the orbitals of
	Neon has 8 valence electrons.	Its orbital-notation electr	on configuration is:
		↑↓ ↑↓ ↑↓↑↓↑↓	
		$\frac{1}{15}$ $\frac{1}{25}$ $\frac{1}{20}$ $\frac{1}{20}$	
		10 L0 Lp	
	Its Lewis dot diagram is :Ne:		

Valence Electrons

Deta	ils	
νια	1113	

Big Ideas

Homework Problems

Fill in the chart below. Use the first row as an example.

Element	Electron Configuration	Group #	Valence Electrons	Lewis Dot	Nearest Noble Gas	Charge of Ion
N	[He] 2s ² 2p ³	15	5	•N:	Ne	-3
0						
Na						
Р						
Ar						
AI						
Br						
В						
Са						
С						
Cl						

Pig Ideas	Introduction: Periodicity	Page: 233
	Introduction: Periodic	i+v
		ity
	Topics covered in this chapter:	
	Development of the Periodic Table	234
	Regions of the Periodic Table	239
	Ionization Energy	244
	Electronegativity	248
	Atomic & Ionic Radius	251
	Standards addressed in this chapter:	
	Massachusetts Curriculum Frameworks & Science Practic	es (2016):
	main group elements, including ionization energy an and ions, based on the patterns of electrons in the o of each element. Use the patterns of valence electro charge, and Coulomb's law to explain and predict ge ionization energies, relative sizes of atoms and ions, elements.	d relative sizes of atoms utermost energy level on configurations, core neral trends in and reactivity of pure

Development of the Periodic Table

Unit: Periodicity

Details

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Describe how our understanding of groups of elements with similar properties developed into the modern periodic table.
- List scientists who contributed to our understanding of periodicity and their contributions.

Success Criteria:

• Descriptions include specific developments and prior developments that they supplanted.

Language Objectives:

• Describe how the modern periodic table developed over time.

Notes:

The Fifth Element

In ancient times, the world was believed to be made of four elements: earth, air, fire, and water. Every substance on Earth was thought to be made of one of these four elements.

When sulfur was discovered, it could be detected in the presence of each of the other four elements. For this reason, philosophers eventually decided that sulfur must be a fifth element.

Chemical Symbols

Details

Big Ideas

By the early 1800s, more than twenty separate elements were known. Swedish chemist Jöns Jacob Berzelius developed a system of notation in which each element was given a one- or two-letter symbol based on its Latin name. This is the system in use today, though newer symbols are based on the elements' names in English rather than Latin. Some of Berzelius's names and symbols included:

Symbol	Latin Name	English Name	Symbol	Latin Name	English Name
Ag	argentum	silver	Na	natrium	sodium
Au	aurum	gold	Pb	plumbum	lead
Cu	cuprum	copper	Sb	stibium	antimony
Fe	ferum	iron	Sn	stannum	tin
Hg	hydrargyrum	mercury	W	wolfram	tungsten
К	kalium	potassium			

Berzelius also developed a system of notation for chemical compounds with the number of atoms of an element denoted as a superscript (*e.g.*, H^2O). Later in the 19th century, the superscript was changed to a subscript, in order to avoid confusing the number of atoms in a molecule with exponents, electrical charges, *etc.*, so H^2O became H_2O .

Early Classifications of the Elements

<u>Johnann Wolfgang Döbereiner</u>: German chemist; classified elements into triads (groups of three) with similar properties (1817).

<u>John Newlands</u>: British chemist; arranged the elements in order of increasing atomic mass. Proposed the *law of octaves* (1829).

<u>law of octaves</u>: when the elements are arranged by increasing atomic mass, every 8th element has similar properties. This worked for the first 20 elements.

n:		
Big	ideas	

Valence Numbers

Details

<u>valence number</u>: the "combining power" of an element, meaning the number of atoms that would need to combine with that element in order to "satisfy" it. Initially, the valence number was the largest number of atoms of oxygen that could form a compound with the element. Note that valence numbers were in use in the early 1800s, almost a century before electrons were discovered.

Valence Number	Metals	Valence Number	Metals	Valence Number	Metals
I.	K, Na, Li, Ca, Ba, Sr, Mg	111.	Cu, Pb, Sb, Bi, U, Ti, Ce, Te	V.	Hg, Ag, Au, Pt, Pd, Rh, Os, Ir
١١.	Mn, Fe, Zn, Sn, Cd, Co, Ni	IV.	As, Mo, Cr, V, W, Ta	VI.	Be, Zr, Y, Th, Al, Si

We now define the valence number to be the number of electrons that an atom has in its outer (valent) shell.

The Periodic Table

<u>Julius Lothar Meyer</u>: German chemist. Published the first periodic table (1864), with 28 elements arranged in order of increasing atomic mass and grouped according to valence numbers.

<u>Dmitri Mendeleev</u>: Russian chemist, considered the author of the modern periodic table. Mendeleev's table was published in 1869, with elements arranged by increasing atomic mass, and grouped in columns by similar chemical & physical properties (including valence number).

Development of the Periodic Table

Page: 237 Unit: Periodicity

Big Ideas

Mendeleev's Periodic Table

Details

(Highlights within vertical columns indicate Döbereiner's triads.)

				Gro	oup			
Period	1	II	III	IV	V	VI	VII	VIII
1	H=1							
2	Li=7	Be=9.4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27.3	Si=28	P=31	S=32	Cl=35.5	
4	K=39	Ca=40	?=44	Ti=48	V=51	Cr=52	Mn=55	Fe = 56; Co = 59; Ni = 59
5	Cu=63	Zn=65	?=68	?=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	?=100	Ru=104; Rh=104; Pd=106
7	Ag=108	Cd=112	ln=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	Pi=138	?Ce=140				
9								
10			?Er=178	?La=180	Ta=182	W=184		Os=195; lr=197; Pt=198
11	Au=199	Hg=200	TI=204	Pb=207	Bi=208			
12				Th=231		U=240		

Mendeleev correctly predicted the existence and chemical and physical properties of undiscovered elements gallium ("eka-aluminum", labeled "?=68" above), and germanium ("eka-silicon", labeled "?=72" above). Mendeleev's periodic table gained significant credibility when gallium and germanium were discovered (during Mendeleev's lifetime) and were found to have the properties that he predicted.

Mendeleev's group numbers were chosen first based on the elements' chemical and physical properties, and second by increasing atomic mass. Tellurium (Te) was later found to have an average atomic mass of 128 (heavier than iodine, which was labeled "J" on Mendeleev's table), but Mendeleev kept the elements where they were, because tellurium is more like the other elements in group 6, and iodine is more like the other elements in group 7.

<u>periodic law</u>: when the elements are arranged in order of increasing atomic number, their properties repeat in regular intervals (periods). (This is the modern version of the law of octaves.) The periodic table is arranged so that each row represents one of these periods.

Development of the Periodic Table

Big Ideas	Details Unit: Periodicit	y
	<u>Henry Moseley</u> : British chemist; rearranged the elements by increasing nuclear charge (atomic number) instead of atomic mass (1913). This arrangement resulted in elements with similar properties falling into groups without the exceptions (such as the positions of iodine <i>vs.</i> tellurium) that occurred in Mendeleev's table. The modern periodic table is Moseley's table, extended to include elements that have been discovered since his death.	

Regions of the Periodic Table

Unit: Periodicity

Details

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

- Identify regions of the periodic table by name.
- Describe the properties of different groups (families) of elements.

Success Criteria:

- Regions of the periodic table are identified correctly.
- Descriptions of properties are correct.

Tier 2 Vocabulary: period, group, family

Language Objectives:

• Name each of the regions of the periodic table.

Notes:

<u>period</u>: a row of the periodic table. Properties of the elements are *periodic*, meaning that they repeat after a specific interval. Elements in the same period have their highest energy electrons in the same principal energy level.

<u>group</u> (family): a column of the periodic table. Elements in the same group have the same number of valence electrons, and therefore have similar chemical and physical properties.

<u>diatomic elements</u>: elements whose natural state is a molecule that has two atoms of the element. There are seven diatomic elements: H₂, N₂, O₂, F₂, Cl₂, Br₂, and I₂.

Regions of the Periodic Table



	Regions of the Perioc	Page: 24	
Big Ideas	Details		Unit: Periodici
	 properties of metals: shiny high density good conductors of heat & electricity malleable & ductile (can be reshaped by hammering, bending and stretching) 	 high melting most have a electrons tend to form 	g & boiling points 3 or fewer valence n positive ions
	 properties of non-metals: dull low density poor conductors of heat & electricity brittle 	 low melting most have 4 electrons tend to forr 	& boiling points for more valence n negative ions
	properties of metalloids: Metalloids can have properties "in betw metals and others like non-metals.	/een," or can have	some properties like

Regions of the Periodic Table



Regions of the Periodic Table

Big Ideas	Details	Unit: Periodicity
	transition metals: elements in the center sect table.	ion (groups 3–12) of the periodic
	 have a partially-filled <i>d</i> sub-level form colored ions when dissolved in water "officially" have 2 valence electrons, but can shift electrons into and out of <i>s</i> and <i>d</i> sub- levels. Often form more than one kind of ion. 	 transition metals with several unpaired electrons in their <i>d</i> or <i>f</i> sub-levels are paramagnetic (are attracted to a magnet). most are shiny, hard metals with high melting & boiling points
	section below the rest of the table.)	ock" of the periodic table. (The "extra"
	 are part of the transition metals 	
	 have a partially-filled f sub-level 	
	 officially have 2 valence electrons, but c sub-levels. Usually form ions with +3 ch 	an shift electrons between <i>s</i> , <i>d</i> , and <i>f</i> arges.
	• are rare	
	noble gases: elements in group 18 (VIIIA) of the	ne periodic table.
	 8 valence electrons (except for He which 	n has 2)—full valence shells
	 do not form ions 	
	 do not react with other compounds 	
	• gases	
	 extremely low melting & boiling points. a solid even at absolute zero, except at e 	(In fact, helium cannot be made into extremely high pressures.)
	halogens: elements in group 17 (VIIA) of the p	periodic table.
	 7 valence electrons (form −1 ions) 	
	reactive	
	 diatomic (atoms in pairs) in their natura 	l state: F ₂ , Cl ₂ , Br ₂ , I ₂
	 low melting & boiling points. (F & Cl are and I is a solid, but will melt in your hand 	e gases at room temp; Br is a liquid, d.)
	 form salts that are soluble in water (exc soluble in water.) 	ept for fluorine—fluoride salts are not
	Homew	ork
	Color and label the regions of the periodic tab elements and data).	le on an actual periodic table (with

Ionization Energy

Unit: Periodicity

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Rank elements according to ionization energy based on their location on the periodic table.

Success Criteria:

• Rankings account for size and effective nuclear charge.

Tier 2 Vocabulary: ionization

Language Objectives:

• Explain why ionization energy increases as you go up and to the right on the periodic table.

Notes:

<u>ionization energy</u>: the amount of energy that it takes to remove an electron from an atom. (This makes it an ion. Ionization energy is literally the amount of energy it takes to make an atom into an ion.) Ionization energy is a measure of how tightly an element holds onto its electrons.



Ionization Energy

Big Ideas	Details		Unit: Periodicity
	The more an element "wants" to favorable it is), the more tightly it the ionization energies.	gain electrons (meaning the more will hold onto its own electrons, a	energetically nd the higher
	 Atoms of elements farther because obtaining a full val obtaining a full valent shell left are more stable if they gaining them. 	to the right hold on to electrons me ent shell by gaining electrons is mo by losing them. Similarly, element attain a full valent shell by losing el	ore tightly, ore stable than s farther to the lectrons than by
	 Atoms of elements higher u because they have fewer en over. Also, because the ato nucleus, which means the f stronger. 	ip within a group hold on to electro nergy levels & sublevels to spread t om has fewer levels, the electrons a orce of attraction by the positive cl	ons more tightly he electrons are closer to the harge is
	 Noble gases and ions that h have the highest ionization electron configurations. 	ave the same electron configuration energies, because they have the m	on as a noble gas nost stable
	Helium has the highest ioni	zation energy. Francium has the lo	west.
	Ionization energy is calculated by element to lose an electron and b	measuring how much heat it takes become a positive ion.	s to get an
	<u>1st ionization energy</u> : the amount neutral atom, to make a +1 io	t of energy it takes to remove one on.	electron from a
	2 nd ionization energy: the amoun make a +2 ion from a +1 ion.	t of energy it takes to remove a sec	cond electron to
	3 rd , 4 th ionization energy, etc.: the electron to make a +3 ion fro +4 ion, etc.	e amount of energy it takes to rem m a +2 ion, to remove a fourth elec	ove a third stron to make a

Ionization Energy

Big Ideas Details

Consider the following table of ionization energies.

Element	1 st Ionization Energy (kJ/mol)	2 nd Ionization Energy (kJ/mol)	3 rd Ionization Energy (kJ/mol)
	(neutral atom)	(+1 ion)	(+2 ion)
Ne	2081	3952	6122
Na	496	4562	6912
Mg	738	1451	7733

Notice that there is a large jump in the ionization energy once the atom or ion has a full valent shell (the same electron configuration as a noble gas, or a "noble gas core").

In the above table, neon (Ne) has the largest 1st ionization energy, because it is a noble gas and it already has a full valent shell.

Sodium (Na) has the smallest 1^{st} ionization energy. This is because removing one electron will give it a "noble gas core". However, sodium has the largest 2^{nd} ionization energy, because the +1 ion has a full valent shell, and is therefore more stable.

Magnesium (Mg) has the lowest 2^{nd} ionization energy, because removing that second electron will give it a noble gas core. However, because the +2 ion has a noble gas core, Mg has the largest 3^{rd} ionization energy.

First ionization energies for all elements are listed in your Chemistry Reference Tables in "Table Z. Selected Properties of the Elements," which begins on page 516.

Big Ideas	Details	Unit: Periodicity
		Homework Problems
	For each pair	of elements:
	 Answe energy 	r the question about which element has the higher or lower ionization .
	 State t that yo 	he direction(s) on the periodic table (up <i>vs.</i> down and/or left <i>vs.</i> right) ou based your choice on.
	 Explain the dif 	n <i>why</i> moving that direction (up <i>vs.</i> down and/or left <i>vs.</i> right) caused ference in ionization energy.
	1. Which Direct	n element has a <i>higher</i> first ionization energy: Na or Al ? :ion(s):
	Expla	nation:
	2. Which Direct	n element has a <i>higher</i> first ionization energy: Mg or Ca ? :ion(s):
	Expla	nation:
	3. Which Direct Explan	n element has a <i>lower</i> second ionization energy: K or Ca ? tion(s): nation:
	4. Which Direct Explan	n element has a <i>lower</i> second ionization energy: Sr or Ba ? :ion(s): nation:
	5. Which Direct Explan	n element has a <i>higher</i> first ionization energy: K or B ? tion(s): nation:
	6. Which Direct Explan	n element has a <i>lower</i> first ionization energy: Rb or P ? tion(s): nation:

Electronegativity

Unit: Periodicity

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Rank elements according to electronegativity energy based on their location on the periodic table.

Success Criteria:

• Rankings account for electron configuration and size.

Tier 2 Vocabulary:

Language Objectives:

• Explain why electronegativity increases as you go up and to the right on the periodic table.

Notes:

<u>electronegativity</u>: the tendency of an atom to attract electrons.



Flectronegativity

		-									1 490	
g Ideas	Details									Unit	: Perio	dic
	Elements that "want" electrons (for which attracting electrons is energetically favorable) pull harder, which makes them more electronegative.											
	 Atoms of elem gases) pull har gaining electro positive ion (for 	ents der, k ons. E ormed	farthe becau ileme l by lo	er to t se it's nts to osing e	he rig easie the le electre	ht on st for eft pu ons) is	the po them II the I s more	eriodi to ge least e stab	ic tabl et a fu hard, lle tha	e (excep Il valence because in the ne	t for no e shell the gative	ob by io
	 Atoms of elem they have fewe makes it harde 	ents er ene er for	highe ergy le them	r up w evels a to be	vithin & sub stabl	a grou levels e with	up or t to spi n only	family read t their	y pull the el existi	harder b ectrons c ng electr	ecause over, w ons.	∍ vhi
	 Noble gases do electrons, nor 	o not "pusł	have n" the	electr m aw	onega ay.	ativitie	es, be	cause	they	neither '	'pull"	
	 Fluorine is the Francium is the the Pauling sca 	• Fluorine is the most electronegative element (3.98 on the Pauling scale). Francium is the least electronegative (or most electropositive) element (0.7 of the Pauling scale).										
	Pauling originally ass steps of 0.5 across p	ignec eriod	l arbit 2 on ⁻	the pe	electro eriodi	onega c table	tivity e:	numt	oers tl	nat increa	ased in	ı
		Li	Ве	В	С	Ν	0	F	Ne			
		1.0	1.5	2.0	2.5	3.0	3.5	4.0	-			
	Electronegativity is c to break bonds betw	urren een a	itly ca itoms	lculat of dif	ed by feren	comp t elen	baring nents.	the a	imour	nt of ene	rgy it t	ak
	Now that we are able to make more precise measurements, the current values of the electronegativities for these elements are close to, but not exactly equal to Pauling's original numbers.											
	Electronegativities for all elements are listed in your Chemistry Reference Tables in "Table Z. Selected Properties of the Elements," which begins on page 516.											
			nu 200	d/or a	dditio	naln	ator					

Big Ideas	Details Unit: Periodicity
	Homework Problems
	For each pair of elements:
	 Answer the question about which element has the higher or lower electronegativity.
	 State the direction(s) on the periodic table (up vs. down and/or left vs. right) that you based your choice on.
	• Explain <i>why</i> moving that direction (up <i>vs.</i> down and/or left <i>vs.</i> right) caused the difference in electronegativity.
	 Which element is <i>more</i> electronegative: Li or K ? Direction(s):
	Explanation:
	 Which element is <i>more</i> electronegative: P or Cl ? Direction(s):
	Explanation:
	 Which element is <i>more</i> electronegative: Al or N ? Direction(s): Explanation:
	 4. Which element has a <i>lower</i> second ionization energy: Mg or Rb ? Direction(s): Explanation:

Atomic & Ionic Radius

Unit: Periodicity

MA Curriculum Frameworks (2016): HS-PS1-1

Mastery Objective(s): (Students will be able to...)

• Rank elements according to atomic or ionic radius based on their charge, electron configuration and location on the periodic table.

Success Criteria:

• Rankings account for charge, electron configuration and size.

Tier 2 Vocabulary: radius

Language Objectives:

• Explain why atomic and ionic radius decrease as you go up and to the right on the periodic table.

Notes:

<u>atomic radius</u>: the average distance from the nucleus to the outermost electrons in an atom. The atomic radius is a measure of the "size" of the atom.



Atomic & Ionic Radius

Big Ideas	Details	Unit: Periodicity						
	<u>shielding</u> : when electrons in the lower energy levels (closer to the n (cancel) some of the nucleus's positive charge. This causes the to be held less tightly by the nuclear charge, and the atom gets	ucleus) shield outer electrons arger.						
	 Atoms of elements get larger as you move down a column, be adds a new energy level, and each new level is farther out from than the previous level. Also, the inner electrons shield some the nucleus, which means that the nucleus holds onto the out tightly. 	cause each row m the nucleus of the charge in er electrons less						
	• Atoms of elements get smaller as you move to the right withir This is because the amount of unshielded positive charge from greater, which pulls the electrons closer.	n the same row. In the nucleus is						
	Covalent atomic radii for all elements are listed in your Chemistry Re in "Table Z. Selected Properties of the Elements," which begins on p	eference Tables Dage 516.						
	Use this space for summary and/or additional notes:							
Big Ideas	Details Unit: Period							
-----------	---	--	--	---	---	--	--	--
			Ionic R	adius				
	Because electror the ator	Because most of the space that an atom takes up is outside the nucleus (where the electrons are), changing the number of electrons (making an ion) changes the size of the atom.						
	 If present at bread bread th 	you take av ositive char com of the r ecause the ne electrons you add ele narge are la eutral elemo	way electrons, the ion get ge are smaller than the n neutral element with the positive ions have more u closer. ectrons, the ion gets large rger than the neutral ato ent with the same numbe	ts smaller. This me eutral atom and al same number of e unshielded positive er. This means ion m and also larger t er of electrons.	eans ions with a so smaller than an electrons. This is e charge, which pulls s with a negative than an atom of the			
		Element	Covalent radius (of neutral atom) (Å)	Charge of ion	Radius of ion (Å)			
		0	0.75	-2	1.4			
		F	0.73	-1	1.33			
		Ne	0.72	0	—			
		Na	1.54	+1	1.02			
		Mg	1.36	+2	0.72			
		AI	1.18	+3	0.54			

Atomic & Ionic Radius

Big Ideas	Details Unit: Periodicity
	Homework Problems
	For each pair of elements:
	 Answer the question about which element has the larger or smaller atomic or ionic radius.
	 State the direction(s) on the periodic table (up vs. down and/or left vs. right) that you based your choice on.
	 Explain why moving that direction (up vs. down and/or left vs. right) caused the difference in atomic radius.
	1. Which atom is the <i>smallest</i> : Be, C, or F ?
	Direction(s):
	Explanation:
	2. Which atom is the <i>smallest</i> : N, P, or Sb ?
	Direction(s):
	Explanation:
	3. Which atom is the <i>largest</i> : Ba, Ga, or N?
	Direction(s):
	Explanation:
	4. Which ion is the <i>largest</i> : P^{3-} , S^{2-} , or Cl^- ?
	Direction(s):
	Explanation:
	5. Which ion is the <i>smallest</i> : Na^+ , Mg^{2+} , or Al^{3+} ?
	Direction(s):
	Explanation:
	6. Which of the following is the <i>smallest</i> : O^{2-} , Ne, or Mg^{2+} ?
	Direction(s):
	Explanation:

Introduction: Nomenclature & Formulas

Unit: Nomenclature & Formulas

Details

Topics covered in this chapter:

Bonding	256
Chemical Formulas	258
Balancing Charges	259
Polyatomic lons	263
Naming Ionic Compounds	
Naming Oxyanions	270
Naming Acids	273
Naming Molecular (Covalent) Compounds	275

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

HS-PS2-6 Communicate scientific and technical information about the molecular-level structures of polymers, ionic compounds, acids and bases, and metals to justify why these are useful in the functioning of designed materials.

		Bonding	Page: 256			
Big Ideas	Details		Unit: Nomenclature & Formulas			
		Bondin	g			
	Unit: Nomenclature 8	& Formulas				
	MA Curriculum Fran	neworks (2016): HS-PS2	-6			
	Mastery Objective(s): (Students will be able	e to)			
	 Explain how ato 	oms bond together to form	compounds.			
	 Identify differer 	nt types of chemical bonds.				
	Success Criteria:					
	 Explanations ac 	count for sharing or transfe	er of electrons.			
	Tier 2 Vocabulary:	bond				
	Language Objective	s:				
	 Explain what hat 	ppens with electrons in or	der to form chemical bonds.			
	Notes:					
	bonding: any joining together of atoms or molecules					
	chemical bond or intramolecular bond: a strong bond between atoms or individual ions, resulting from the sharing or transfer of electrons					
	intermolecular bond: molecules of a liq in the section on f	a weak bond between mo uid or solid together. (We "Intermolecular Forces" on	lecules or ions, which holds the will study these in more detail later page 313.)			
	ion: an atom or group lost electrons.	o of atoms that has a charg	e, because it has either gained or			
	Use this space for sun	nmary and/or additional no	otes:			

	Bonding	Page: 257					
Big Ideas	Details Unit: Nomenc	lature & Formulas					
	Types of Chemical Bonds	Types of Chemical Bonds					
	ionic bond: when a positive ion and a negative ion are held together electrical attraction of their charges.	er by the					
	 ionic bonds occur between ions, usually between a metal io metal ion. 	on and a non-					
	 the positive ion (<u>cation</u>) is always either the ion of a metal polyatomic ion. 	or a positive					
	 the negative ion (anion) is always either the ion of a nonmer polyatomic ion. 	etal or a negative					
	• the difference between the electronegativity of the nonme electronegativity of the metal (ΔX) is usually \geq 1.7. (This w further in the section on "Intermolecular Forces" on page 3	tal and the ill be addressed 313.)					
	<u>covalent bond</u> : when two atoms form a bond by sharing ("co-") the valent") electrons.	eir valence ("-					
	 covalent bonds occur only between non-metals 						
	• the electronegativity difference ($\Delta \chi$) between the two non < 1.7	metals is usually					
	<u>metallic bond</u> : when atoms in a metal form a network of positive in held electrons.	ons and loosely					
	 metallic bonds, as the name suggests, occur only between metallic bonds are often described as a "sea of electrons" is valence electrons can move easily from one atom to anoth metals conduct electricity—electricity is simply the flow of 	metals because the er. This is how electrons.					
	Lise this space for summary and/or additional notes:						

Chemical Formulas

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to...)

• Write chemical formulas for ionic compounds when the number of atoms of each element is given.

Success Criteria:

- Elements are listed in the correct order.
- Subscripts give the correct number for each element.

Tier 2 Vocabulary: formula

Language Objectives:

• Explain what the subscripts mean in a chemical formula.

Notes:

<u>chemical formula</u>: a formula that describes a compound by listing how many of each element it's made of.

Some examples:

- Fe₂O₃ has 2 Fe (iron) atoms and 3 O (oxygen) atoms.
- CaCl₂ has 1 Ca (calcium) atom and 2 Cl (chlorine) atoms.
- C₂₁H₃₀O₂ has 21 C (carbon) atoms, 30 H (hydrogen) atoms, and 2 O (oxygen) atoms.

Elements in a chemical formula are listed with metals first, then non-metals, and almost always in order by increasing electronegativity: the <u>least</u> electronegative element is listed first, and the most electronegative one is listed last. (Exceptions are organic compounds and acids.)

(Note: the variable χ is usually used for electronegativity.)

For example: a compound made from Mg²⁺ ions (χ_{Mg} = 1.31) and Cl⁻ ions (χ_{Cl} = 3.16) would be MgCl₂, not Cl₂Mg.

Balancing Charges

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to...)

• Write chemical formulas for ionic compounds.

Success Criteria:

• Subscripts are chosen so that positive and negative charges are balanced (equal).

Tier 2 Vocabulary: bond, charge

Language Objectives:

• Explain the process and necessity of balancing charges.

Notes:

Details

If you have an ionic compound (a compound made of positive and negative ions), the positive and negative charges will attract each other (because opposite charges attract). This will continue to happen until the total amount of positive charge equals the total amount of negative charge, and there is no more attraction. When this happens in chemistry, we say that the charges are <u>balanced</u>.





4. Now the group is positive, so it will attract another negative O²⁻ ion, giving us:



- 5. Finally, all of the positive and negative charges have exactly balanced, and the compound has no net charge. Now it doesn't attract any more positive or negative ions.
- 6. To balance the charges, we needed 2 AI^{3+} ions and 3 O^{2-} ions, which means the formula for this compound is $(AI^{3+})_2(O^{2-})_3$ or simply AI_2O_3 .

Shortcuts for Balancing Charges

Find the L.C.M.

Details

In an ionic compound, *the total positive and total negative charge will always be equal*, and will be the least common multiple (L.C.M.) of the charges of the positive and negative ions.

In the compound made from aluminum and oxygen the charges of the ions are +3 (for Al), and -2 (for O). The LCM of 3 and 2 is 6, which means the total positive charge in the formula will be +6, and the total negative charge in the formula will be -6.

To get +6, we need 2 Al³⁺ ions, and to get -6 we need 3 O²⁻ ions. Thus, the formula is once again (Al³⁺)₂(O²⁻)₃ or simply Al₂O₃.

Cross the Charges and Reduce to Lowest Terms

Often, you can "cross the charges" to get the formula:

Al³⁺₂ 2²⁻

This always gives a correct ratio, although the ratio may not be in lowest terms. (For example, crossing the charges for Fe^{2+} and O^{2-} would give the compound Fe_2O_2 , but the correct formula should be reduced to FeO.)

Use this space for summary and/or additional notes:

Big Ideas

In the chart below,

Details

Big Ideas

• Look up and add the appropriate charges to the cation and anion (if the charges are not already labeled).

Homework Problems

• Balance the charges and write the formula of the resulting compound.

Cation	Anion	Chemical Formula
NH_4^+	(PO ₄) ³⁻	(NH ₄) ₃ PO ₄
Sr	S	
Na	CI	
Са	Br	
к	О	
Cu⁺	CI	
Cu ²⁺	CI	
Mg	S	
Ва	Ρ	
Cr ⁶⁺	0	

Polyatomic Ions

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to...)

• Write chemical formulas that include polyatomic ions.

Success Criteria:

- Subscripts are chosen so that positive and negative charges are balanced (equal).
- Formulas for polyatomic ions are in parentheses if more than one is needed.

Tier 2 Vocabulary: bond, charge

Language Objectives:

• Explain the process and necessity of balancing charges.

Notes:

<u>polyatomic ion</u>: a group of atoms that are bonded to each other that behave chemically like a single ion. A polyatomic ion always has a specific name, chemical formula, and charge.

For example: the sulfate ion has the chemical formula SO_4^{2-} . It is made of one sulfur atom and 4 oxygen atoms. Chemically, it behaves like a single atom with a -2 charge.

The formula of a polyatomic ion never changes!

I.e., the sulfate ion is *always* SO_4^{2-} , and the 4 is an important part of the formula. If you wrote SO_2^{2-} instead, you would be talking about the hyposulfite ion instead of the sulfate ion—a different polyatomic ion with different chemical properties.

If a compound contains a polyatomic ion, you write the formula for the polyatomic ion, *including the subscript numbers*, in the place where the ion goes. For example, a compound with Na^+ and $SO_4^{2^-}$ would simply be Na_2SO_4 .

Balancing Charges with Polyatomic Ions

If you need more than one of a polyatomic ion in a chemical formula, put the entire polyatomic ion, *including any subscript numbers*, in parentheses, and put the number that tells how many ions you need outside the parentheses.

For example, to balance the compound made from AI^{3+} and SO_4^{2-} , you need 2 AI^{3+} ions and 3 SO_4^{2-} ions. The formula is:

 $AI_2(SO_4)_3$

Note: there are positive and negative polyatomic ions. A compound can have either, neither, or both kinds. For example, if you had a compound made from the positive ion ammonium (NH_4^+) and the negative ion sulfate (SO_4^{2-}), the compound would have the formula:

 $(NH_4)_2 SO_4$

Determining the Number of Atoms in a Formula

The subscripts tell you how many you have of *whatever came immediately before the subscript*. If the thing before the subscript is an element, as in $CaCl_2$, the 2 tells us that we have 2 Cl atoms. There's no subscript after Ca, so this means we have only 1 Ca atom.

If the thing before the subscript is parentheses, as in $Al_3(SO_4)_2$, the 3 tells us that we have 3 Al atoms, the 2 outside the parentheses tells us that we have 2 entire SO_4 ions. This means we really have 2 atoms of S and 2 × 4 = 8 atoms of O.

Sample Problem:

How many hydrogen atoms are in the compound $(NH_4)_2 HPO_4$?

We have $2 \times 4 = 8$ from the two NH₄ ions, plus 1 from the HPO₄ ion, giving us a total of 9 hydrogen atoms.

Use this space for summary and/or additional notes:

Big Ideas

Details

Polyatomic Ions

Big Ideas

Details

Table of Polyatomic Ions							
ion	formul a	ion	formula	ion	formula		
americyl	AmO ₂ ²⁺	acetate	CH ₃ COO - tetraborate		B ₄ O ₇ ²⁻		
carbonyl	CO ₂ ²⁺	amide	$\rm NH_2^-$	carbide	C ₂ ²⁻		
thiocarbonyl	CS ₂ ²⁺	hydroxylamide	NHOH⁻	carbonate	CO3 ²⁻		
chromyl	CrO ₂ ²⁺	azide	N₃ [−]	chromate	CrO4 ²⁻		
neptunyl	NpO ₂ ²⁻	hydrazide	$N_2H_3^-$	dichromate	$Cr_2O_7^{2-}$		
plutoryl	PuO ₂ ²⁺	bromate	BrO₃ [−]	imide	NH ²⁻		
seleninyl	SeO ²⁺	chlorate	ClO₃ [−]	molybdate	MoO4 ²⁻		
selenoyl	SeO ₂ ²⁺	cyanide	CN⁻	peroxide	O ₂ ²⁻		
thionyl/sulfinyl	SO ²⁺	cyanate	OCN⁻	oxalate	$C_2 O_4^{2-}$		
sulfonyl/sulfuryl	SO ₂ ²⁺	thiocyanate SCN ⁻ phthalate		phthalate	$C_8H_4O_4{}^2$		
uranyl	UO ²⁺	selenocyanate	SeCN⁻	selenite	SeO4 ²⁻		
vanadyl	VO ²⁺	tellurocyanate	CH₃S⁻	silicate	SiO ₃ ²⁻		
ammonium	NH_4^+	hydroxide	OH⁻	sulfate	SO4 ²⁻		
hydronium	H₃O⁺	iodate	1O ₃ -	thiosulfate	$S_2O_3^{2-}$		
iodyl	IO_2^+	methanolate	CH₃O⁻	dithionate	$S_2O_4^{2-}$		
nitrosyl	NO⁺	methanethiolate	CH₃S⁻	silicate	SiO ₃ ²⁻		
thionitrosyl	NS⁺	ethanolate	C₂H₅O⁻	borate	BO ₃ ³⁻		
phosphoryl	PO⁺	permanganate	MnO₄⁻	arsenate	AsO4 ³⁻		
thiophosphoryl	PS⁺	nitrate	NO₃ [−]	phosphate	PO4 ³⁻		
phosphor	PO_2^+	superoxide	O ₂ ⁻	orthosilicate	SiO4 ⁴⁻		

Naming Ionic Compounds

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to...)

- Write names for ionic compounds using the stock system, including Roman numerals where appropriate.
- Write formulas for ionic compounds based on their names.

Success Criteria:

Details

Big Ideas

- Compound names contain the correct cation (including a Roman numeral if necessary) and anion in the correct order (cation first, then anion).
- Chemical formulas have correctly balanced charges.
- Chemical formulas have polyatomic ions in parentheses when necessary.

Tier 2 Vocabulary: compound

Language Objectives:

- Explain when a cation does or does not need a Roman numeral.
- Explain the relationship between the charges of the ions and the subscripts in the formula.

Notes:

ionic compound: a compound made out of the ions of a metal and a nonmetal

cation: an ion with a positive charge, such as Na⁺ or Ca²⁺

anion: an ion with a negative charge, such as Cl⁻ or S²⁻

Naming the Anion

- If an anion is a single element, the name of the ion is the name of the element with the ending changed to "ide". For example, the Cl⁻ ion is made from chlorine, so it is called "chloride". The O²⁻ ion is made from oxygen, so it is called oxide.
- If the anion is a polyatomic ion, its name is just the name of the polyatomic ion. For example, the NO₃[−] ion is named "nitrate".

Details

Naming the Cation

- If the cation is a single element that has only one possible charge, the name of the cation is the name of the element. For example, the K⁺ ion is simply named "potassium", and the Ca²⁺ ion is simply named "calcium".
- If the element can have more than one possible charge, the name of the cation is the name of the element followed by a Roman numeral, indicating the charge, in parentheses. For example, chromium can make cations with three different charges:

Formula of Cation	Name of Cation	
Cr ²⁺	chromium (II)	
Cr³⁺	chromium (III)	
Cr ⁶⁺	chromium (VI)	

• If the cation is a polyatomic ion, its name is the name of the polyatomic ion. For example, the NH₄⁺ ion is named "ammonium".

Naming the Compound

stock system: a system of naming compounds by naming the ions that they're made of. The cation (positive ion) is always listed first and the anion (negative ion) is always listed last.

Examples:

Formula	Cation		А	nion	Name of Compound
NaCl	Na⁺	sodium	Cl⁻	chloride	sodium chloride
CaBr ₂	Ca ²⁺	calcium	Br⁻	bromide	calcium bromide
Fe_2O_3	Fe ³⁺	iron (III)	0 ²⁻	oxide	iron (III) oxide
FeO	Fe ²⁺	iron (II)	O ²⁻	oxide	iron (II) oxide
K_2SO_4	K+	potassium	SO4 ²⁻	sulfate	potassium sulfate

Notice that the number of atoms in the chemical formula is not represented anywhere in the name of an ionic compound.

Naming Ionic Compounds

Big Ideas	Details Unit: Nomenclature & Formulas
	Writing Formulas from Names
	The name of every ion must have enough information to describe its element(s) and its charge. This means that if you know the name of an ion, you know the formula and the charge.
	To convert the name of a compound to its formula:
	1. Write down the cation and anion, including their charges.
	 Add subscripts to balance the charges. (Don't forget to put polyatomic ions in parentheses when you need more than one!)
	Sample Problems:
	1. What is the chemical formula for calcium phosphate?
	Calcium is Ca ²⁺ and phosphate is PO ₄ ³⁻ . The L.C.M. of the charges is 6. This means we need 3 Ca ²⁺ ions to get to +6, and 2 PO ₄ ³⁻ ions to get to -6 . Therefore, the formula must be:
	Ca ₃ (PO ₄) ₂
	2. What is the formula for nickel (II) chloride?
	Nickel (II) is Ni ²⁺ (remember that the Roman numeral tells us the charge) and chloride is Cl [−] . The L.C.M. of the charges is 2, which means we need one Ni ²⁺ ion to get to +2, and 2 Cl [−] ions to get to −2. Therefore, the formula must be:
	NiCl ₂

In the chart below:

Details

Big Ideas

• Write the element symbol(s) for the cation (positive ion) and anion (negative ion).

Homework Problems

- Determine the charges of both ions.
- Balance the charges and write the formula of the resulting compound.

Cation Name	Cation Formula	Anion Name	Anion Formula	Formula of Compound
ammonium	NH_4^+	phosphate	PO4 ³⁻	(NH ₄) ₃ PO ₄
				SnCl ₂
				FeO
calcium		bromide		
potassium		oxide		
copper (I)		carbonate		
copper (II)		chloride		
magnesium		nitrate		
ammonium		hydroxide		
barium		phosphate		
chromium (VI)		sulfate		

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to...)

- Write names for ionic compounds that contain oxyanions.
- Write chemical formulas for ionic compounds that contain oxyanions.

Success Criteria:

Details

- Compound names contain the correct cation (including a Roman numeral if necessary) and the correct anion (with prefix and suffix that together determine the number of oxygen atoms).
- Chemical formulas have correctly balanced charges.
- Chemical formulas have polyatomic ions in parentheses when necessary.

Tier 2 Vocabulary: compound, anion

Language Objectives:

• Explain what the prefix and suffix tell about the number of oxygens in an anion.

Notes:

oxyanion (or oxoanion): polyatomic anion (negative polyatomic ion) that contains oxygen.

Examples: NO₃ (nitrate), PO₄ (phosphate), SO₃ (sulfite)

	Naming Oxyanions Page: 27				Page: 271		
Big Ideas	Details Unit: Nomenclature & Formula				re & Formulas		
	The names of oxyanions always end in either "ate" or "ite", depending on the <i>relative</i> number of oxygens.						
	Note that most of the "ate" ions have either three or four oxygens, but there is no way to predict how many oxygens any particular "ate" ion will have.						
	 The "ate" ion has its own specific formula and charge. The "ite" ion has one fewer oxygen than the "ate" ion, and the same charge. 						
	Some example	es:					
		"ate" ion	formula	"ite" ion	formula		
		sulfate	SO4 ²⁻	sulfite	SO ₃ ²⁻		
		chlorate	ClO₃⁻	chlorite	ClO2 [−]		
		nitrate	NO₃⁻	nitrite	NO₂ [−]		
		phosphate	PO4 ³⁻	phosphite	PO3 ³⁻		
	If you have on the charge sta • chlorite	e fewer oxygen th ys the same. For = CIO_2^-	an the " example:	ite" ion, add t	the prefix "	hypo". Again,	
	 hypochl 	orite = ClO⁻					
	If you have on the charge sta	e more oxygen the same. For	an the " example:	ate" ion, add	the prefix "	per". Again,	
	 chlorate 	$e = CIO_3^-$					
	 perchlor 	rate = CIO_4^-					

Naming Oxyanions

Page: 272
Unit: Nomenclature & Formulas

Big Ideas	Details	Unit: Nomenclature & Formula		
	Homework Problems			
	Give the name for each of the following formulas for polyatomic ions in "Table Reference Tables on page 512.	g compounds. You may need to look up K. Polyatomic lons" in your Chemistry		
	1. Rb ₂ B ₄ O ₇	5. Cr(CO₂)₃		
	2. Mo₂(MoO₃)₃	6. Na ₂ SeO ₂		
	3. CoSO ₂	7. NaNO₃		
	4. Ca ₂ SiO ₄	8. KNO ₂		
	Give the chemical formula for each of the following compounds.			
	9. sodium nitrate	13. potassium dichromate		
	10. ammonium periodate	14. magnesium thiosulfite		
	11. calcium hypochlorite	15. potassium hypobromite		
	12. barium carbonite	16. sodium phthalate		

Naming Acids

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to ...)

- Write names for inorganic acids.
- Write chemical formulas for inorganic acids.

Success Criteria:

Details

- Compound names contain the name of the anion with the correct prefix and/or suffix, and the word "acid".
- Chemical formulas have correctly balanced charges.
- Chemical formulas have polyatomic ions in parentheses when necessary.

Tier 2 Vocabulary: acid, formula

Language Objectives:

• Explain what the prefix and suffix tell about the type of anion in an acid.

Notes:

<u>acid</u>: a chemical compound that creates hydrogen (H^+) ions in water.

Acids behave somewhat like ionic compounds in which the cation (positive ion) is H^{+} . (We will study acids and bases in detail later in the year.)

Because the cation is always H^+ , the name of the acid is based on the name of the anion (negative ion).

Anion Ends With	Example	Acid Name	Example
ate	nitrate (NO₃⁻)	ic acid	nitric acid (HNO₃)
ite	arsenite (AsO ₃ ³⁻)	ous acid	arsenous acid (H ₃ AsO ₃)
ide	chloride (Cl⁻)	hydroic acid	hydrochloric acid (HCl)

Any prefixes, such as "per-" and "hypo-", are kept:

- periodate is IO₄⁻ so the acid HIO₄ is periodic acid
- hypochlorite is ClO⁻ so the acid HClO is hypochlorous acid.

A stupid mnemonic that some students seem to like for remembering the pair of suffix changes is: "I *ate* something *ic*ky. It m*ite* be a hippopotam*ous*."

Naming Acids

Dotaile	INdii	ing Acius	Page: 274			
	Ц	mework Proh				
Fill in the chart below. Use the first row as an example.						
Chemical Formula	Anion	Anion Name	Acid Name			
HNO ₃	NO₃⁻	nitrate	nitric acid			
H ₂ CO ₃						
HBr						
			acetic acid			
HNO ₂						
			phosphoric acid			
			sulfurous acid			
			hydroïodic acid			
НСІ						
			chloric acid			
HClO ₂						
	Details Fill in the chart b Chemical Formula HNO3 H2CO3 HBr HNO2 HNO2 HNO2 HNO2 HNO2 HNO2 HNO2	IN a monomeno Details Fill in the chart below. Use th Chemical Formula Anion HNO3 NO3 ⁻ H2CO3 I HBr I HNO2 I HO1 I HCI I HCIO2 I	Naming Actos Details Homework Prob Fill in the chart below. Use the first row as an exar Chemical Formula Anion Anion Name HNO3 NO3 nitrate H2CO3 Image: State S			

Use this space for summary and/or additional notes:

_

Naming Molecular (Covalent) Compounds

Unit: Nomenclature & Formulas

MA Curriculum Frameworks (2016): HS-PS2-6

Mastery Objective(s): (Students will be able to...)

- Write names for molecular (covalent) inorganic compounds.
- Write chemical formulas for molecular (covalent) inorganic compounds.

Success Criteria:

Details

- Compound names contain the name of each element with the appropriate prefix (except that a compound cannot start with "mono-").
- Chemical formulas have elements listed in the correct order (increasing electronegativity).
- Chemical formulas have the correct subscript for each element.

Tier 2 Vocabulary: molecule, compound

Language Objectives:

• Explain how prefixes are used with molecular compounds.

Notes:

<u>molecular compound</u> (also known as a <u>covalent compound</u>): a compound made of atoms joined by covalent bonds (shared electrons).

molecule: a set of atoms joined by covalent bonds.

Naming Molecular (Covalent) Compounds

Big Ideas

NPOUNDS Page: 276 Unit: Nomenclature & Formulas

Unlike ionic compounds, covalent compounds have names that give the chemical formula of the molecule. The numbers in the formula are Greek numbers made into prefixes.

No.	Prefix	No.	Prefix	No.	Prefix	No.	Prefix
1	mono- or hen-	10	deca-	100	hecta-	1000	kilia-
2	di- or do-	20	(i)cosa-	200	dicta-	2000	dilia-
3	tri-	30	triaconta-	300	tricta-	3000	trilia-
4	tetra-	40	tetraconta-	400	tetracta-	4000	tetralia-
5	penta-	50	pentaconta-	500	pentacta-	5000	pentalia-
6	hexa-	60	hexaconta-	600	hexacta-	6000	hexalia-
7	hepta-	70	heptaconta-	700	heptacta-	7000	heptalia-
8	octa-	80	octaconta-	800	octacta-	8000	octalia-
9	nona-	90	nonaconta-	900	nonacta-	9000	nonalia-

Rules:

Details

- 1. The first atom in the formula has a number prefix only if the molecule contains more than one of that atom.
- 2. The last atom in the formula always has a number prefix, whether or not there are more than one of that atom.
- 3. For huge numbers of atoms, the number prefixes combine in *reverse* order of place value. (Yes, it's weird.)

For example, the compound $C_{51}S_{22}$ (if it existed) would be called henpentacontacarbon docosasulfide.

In practice, prefixes larger than 10 (deca-) are rarely used.

Naming Molecular (Covalent) Compounds Page: 277

- N₂O has 2 nitrogen atoms ("di-nitrogen") and one oxygen atom ("monooxide," which we elide to "monoxide"), so its name is dinitrogen monoxide.
- H₂O has 2 hydrogen atoms and one oxygen atom, but it is always "water," not "dihydrogen monoxide". The same goes for NH_3 , which is always called "ammonia."

Someone has created a humorous website, http://www.dhmo.org, which attempts to make people aware of the "dangers" of dihydrogen monoxide, or "DHMO."

Use this space for summary and/or additional notes:

Big Ideas

Details

Examples:

Naming Molecular (Covalent) Compounds

Big Ideas	Details Unit: Nomenclature & Formulas		
	Homework Problems		
	Give the name for each of the following covalent compounds, using the prefix system.		
	1. NF ₃	6. N ₂ O ₄	
	2. NO	7. PCl ₃	
	3. NO ₂	8. PCI₅	
	4. B ₂ O ₃	9. SF ₆	
	5. N ₂ O	10. CO ₂	
	Give the chemical formula for each of the fol	lowing compounds.	
	11. phosphorus triïodide	16. boron trichloride	
	12. sulfur dichloride	17. phosphorus pentafluoride	
	13. xenon trioxide	18. diphosphorus trioxide	
	14. dinitrogen tetrafluoride	19. dichlorine heptoxide	
	15. sulfur tetrafluoride	20. carbon tetrachloride	

Big Ideas

Introduction: Covalent Bonding & Molecular

Geometry

Unit: Covalent Bonding & Molecular Geometry

Topics covered in this chapter:

Covalent Bonding & Lewis Structures	280
Charged Atoms in Lewis Structures	286
VSEPR Theory	289
Orbital Hybridization	295

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

HS-PS1-2 Use the periodic table model to predict and design simple reactions that result in two main classes of binary compounds, ionic and molecular. Develop an explanation based on given observational data and the electronegativity model about the relative strengths of ionic or covalent bonds.

Covalent Bonding & Lewis Structures

Unit: Covalent Bonding & Molecular Geometry

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Draw Lewis structures for diatomic molecules, showing bonds and lone pairs.

Success Criteria:

Details

- Lewis structures show the correct number of bonds.
- Lewis structures show the correct number of unpaired electrons, drawn in pairs.

Tier 2 Vocabulary: bond, lone pair

Language Objectives:

• Explain how a Lewis structure represents the electrons in a molecular compound.

Notes:

<u>covalent bond</u>: a chemical bond consisting of one or more pair(s) of shared electrons.

<u>covalent compound</u> (also known as a <u>molecular compound</u>): a compound made of atoms joined by covalent bonds.

- covalent bonding occurs between non-metals
- electrons are shared in pairs—each pair usually contains one electron that came from each atom.

For example, a chlorine atom has seven valence electrons. The Lewis dot structure for chlorine is:

Note that the chlorine atom has seven valence electrons.

Covalent Bonding & Lewis Structures



Covalent Bonding & Lewis Structures Page: 282

Big Ideas	Details Unit: Covalent Bonding & Molecular Geometry
	For another example, oxygen has six unpaired electrons, which means an oxygen atom needs two more electrons to fill its valent shell. This means that each oxygen atom in O_2 will share two electrons:
	$\ddot{0}$
	Just like the example with Cl_2 , each oxygen atom has a full valent shell whenever it has all four of the shared electrons.
	One pair of shared electrons is called a <u>single bond</u> , such as the bond in Cl_2 . The O_2 molecule has two pairs of shared electrons, which is called a <u>double bond</u> . Three pairs of shared electrons would be a <u>triple bond</u> . Note that a double bond is a large electron cloud with four electrons, and a triple bond is an even larger electron cloud with six electrons.
	An atom can make bonds to more than one other atom. For example, oxygen needs two additional electrons. IT doesn't matter where the oxygen atom gets these electrons from—oxygen can just as easily share one electron with each of two different atoms. An example is the H ₂ O molecule:
	H H
	This reasoning can be extended to any non-metal. The rule of thumb is that atoms have to "share one to get one". This means that the number of bonds to any atom will be equal to the number of electrons the atom needs in order to fill its valent shell.
	Because of space constraints and the shapes of the <i>s</i> and <i>p</i> orbitals involved in forming bonds, it is not possible for atoms to make more than a triple bond to a single atom. Also, in a first-year high school chemistry course, we will only consider structures made by electrons in <i>s</i> and <i>p</i> orbitals. This means the structures we will consider in this course will have no more than four total electron clouds. Structures with expanded octets are studied in $AP^{\text{@}}$ Chemistry.
	Use this space for summary and/or additional notes:

Covalent Bonding & Lewis Structures

Page: 283



		Covalent Bonding & Lewis Structures Page: 284			
Big Ideas	Details	Unit: Covalent Bonding & Molecular Geometry			
	Rules for Drawing Lewis Structures				
	1.	Neutral atoms will make the same number of bonds as the number of additional electrons they need to fill their valent shell.			
	2.	No pair of atoms can have more than a triple bond between them.			
	3.	Atoms that make only one bond (<i>i.e.,</i> that need only one more electron) will almost always be on the outside (<i>i.e.,</i> not in the center).			
	4.	The least electronegative atom will almost always be in the center.			
	5.	The total number of bonds in a Lewis structure will usually be:			
		# electrons wanted – actual # electrons			
		2			
	If you c (These page 28	annot draw a correct Lewis structure using neutral atoms, try the following. are explained in the section titled "Charged Atoms in Lewis Structures" on 86.)			
	6.	Try moving one electron (at a time) from the central atom to the most electronegative atom that doesn't yet have a negative charge.			
	7.	If the structure is for a charged ion, add or remove the appropriate number of electrons from your structure.			
		 a. If you need to add electrons, add them to the most electronegative atom first. 			
		 b. If you need to remove electrons, remove them from the least electronegative atom first. 			
	8.	If you can draw more than one valid structure, the one with the lowest maximum charge on an atom is preferred. If you can draw more than one such structure, the structure with the fewest number of charged atoms is preferred.			

Page: 28	85
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Big Ideas	Details	0	Unit: Covalent Bonding & Molecular Geomet
	Homework Problems Draw a correct Lewis structure for each of the following compounds.		
	1.	PI ₃	11. CH ₂ O
	2.	N ₂	12. CO ₂
	3.	H ₂ O	13. CBrClFl
	4.	PBr ₃	14. NF ₃
	5.	SiCl ₄	15. N ₂ H ₂
	6.	HCI	16. IBr
	7.	FCN	17. CH₃OH
	8.	HNO	18. C ₂ H ₄
	9.	PN	19. OF ₂
	10	CH ₃ NH ₂	20. COCl ₂

Unit: Covalent Bonding & Molecular Geometry

Charged Atoms in Lewis Structures

Unit: Covalent Bonding & Molecular Geometry

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Draw Lewis structures in which one or more atoms has a formal charge.

Success Criteria:

Details

- Lewis structures show the correct number of bonds.
- Lewis structures show the correct number of unpaired electrons in the correct locations.
- Individual charges are assigned correctly. (Positive charges on least electronegative atom, negative charges on most electronegative atom, *etc.*)
- Total charge adds up to the correct value.

Tier 2 Vocabulary: bond, charge

Language Objectives:

• Explain how charges are assigned to Lewis structures.

Notes:

Lewis structures show the shape of a molecule and how the atoms share electrons with each other. If a molecule exists, that means it must have a Lewis structure.

If you can't find a way to draw the Lewis structure for a molecule using neutral atoms, you may need to take electrons away from one atom and distribute them to other atoms, creating atoms with charges.

For example, consider the sulfur dioxide (SO_2) molecule. Sulfur and oxygen both have 6 valence electrons and need two bonds. If you draw the following:

sulfur has enough bonds, but each oxygen needs one more electron.

Charged Atoms in Lewis Structures

Big Ideas Details Unit: Covalent Bonding & Molecular Geometry If you move one electron from sulfur to one of the oxygens, you have the following situation: :o°−s[⊕]−o· (Notice that when an atom in a Lewis structure has a charge, the charge is written next to the atom and circled.) One way to think of charged atoms in Lewis structures is to temporarily substitute the charged atom with a neutral atom that has the same number of valence electrons, draw the Lewis structure, and then switch the atoms back. When we take an electron away from S to make it S⁺, the S⁺ atom has the same number of valence electrons as P, so we could temporarily substitute P for S. Similarly, when we add an electron to O to make it O^- , the O^- atom has the same number of electrons as F, so we could temporarily substitute F for O. If we made both of these substitutions, the compound would be OPF. We would place a double bond between P and O, to give the following Lewis structure: F-P=0 Now, we switch the F back to O⁻ and the P back to S⁺, which gives the correct Lewis structure for SO₂: :0[⊕]−S[⊕]=0 If you are drawing a Lewis structure for a charged polyatomic ion, you need to show the overall charge as well as the charges on the individual atoms. To do this, put the entire structure in square brackets and write the overall charge outside the brackets, as in the example to the right: This means that when we write the formula CaCO₃ for the ionic compound calcium carbonate, the actual arrangement of the atoms is:

Big Ideas	Details	Charged Atoms in Lewis St	ructures Page: 288 alent Bonding & Molecular Geometry
0		Homowork Pr	obloms
	Draw a correct Lewis structure for each of the following compounds.		
	1.	PF ₃	9. NH ₂ SeF
	2.	O ₂	10. CH₃OH
	3.	H ₂ S	11. SiO ₂
	4.	AsCl ₃	12. CH₃COOH
	E	CPr	12 DOE
	Э.		13. FOF
	6.	HI	14. SO ₄ ²⁻
	-		15 NO -
	7.	HSCN	15. NO ₃
	8.	SCl ₂	16. PO ₄ ³⁻
VSEPR Theory

Unit: Covalent Bonding & Molecular Geometry

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Identify the VSEPR shapes and bond angles for simple molecules (one central atom).

Success Criteria:

Details

- VSEPR shapes show the correct number of lone pairs in the correct locations.
- VSEPR shapes have the correct bond angles.

Tier 2 Vocabulary: bond, cloud

Language Objectives:

• Explain how repulsion between electron clouds results in VSEPR shapes.

Notes:

<u>Valence Shell Electron Pair Repulsion (VSEPR^{*}) theory</u>: a theory that the shape of a molecule is determined by the repulsion between electrons in the bonds and unshared pairs of the atoms.

The Lewis structure of a molecule represents the structure in 2 dimensions. The VSEPR shape is the 3-dimensional equivalent.

The VSEPR shapes are determined by the following constraints:

Electrons are all negatively charged, so they repel each other. Valence electrons exist in electron clouds, which can be either:

- unshared electrons (in pairs), attached to only one atom
- as part of a covalent bond (shared pair of electrons) between two atoms

The VSEPR shape of the molecule is the shape that occurs when all of these "clouds" of electrons are as far apart as possible.

* VSEPR is pronounced as if it were written "vesper".

VSEPR Theory

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Big Ideas	Details	Unit: Covalent Bonding & Molecular Geometry
	For example, in CH ₄ , the electron cl four bonds to the hydrogen atoms. means they get as far apart as poss we draw the bonds at 90° angles, w in a 2-dimensional drawing:	ouds around carbon are the These electrons repel, which ible. In the Lewis structure, which is as far apart as possible H
	However, the molecule is really 3-d means the bonds are actually equal <i>sphere</i> . This would result in a 3-din with the hydrogens at 109.5° angles atom:	imensional. This ly spaced around a hensional molecule, s around the carbon
	If we described this molecule as a g would be a regular (all edges and an tetrahedron, with the carbon atom hydrogen atoms at the vertices:	eometric shape, it ngles equal) in the center and H
	This means that, according to VSEP	R theory, CH₄ is a "tetrahedral" molecule.

Big Ideas Details Unit: Covalent Bonding & Molecular of Molecular of Now, consider the NH3 molecule. The Lewis structure looks like The "lone pair" of electrons (above the N atom) and the three bonds all repel each other. H H This gives <u>four</u> electron clouds, just like CH4. However, because the lone pair is closer to the nucleus, it repels the other electrons more strongly than bond electrons. This compresses the bond angles slightly, to about 107.5°. M M This time, the shape of the pyramid, but not a regular means the VSEPR shape of NH3 is The shape at right shows the atom downward), plus the "invisible" lone pair of electrons above. Molecule is therefore "bent". Because the lon are closer to the nucleus, they repel a little more strongly than bond electrons. The StPR shape of the pyramid, but not a regular means the VSEPR shape of NH3 is M M M The StPR shape of the H2O molecule is therefore "bent". Because the lon are closer to the nucleus, they repel a little more strongly than bond electrons. The VSEPR shape of the H2O molecule is therefore "bent". Because the lon are closer to the nucleus, they repel a little more strongly than bond electrons.	Page: 291
Now, consider the NH ₃ molecule. The Lewis structure looks like this: The "lone pair" of electrons (above the N atom) and the three bonds all repel each other. This gives <u>four</u> electron clouds, just like CH ₄ . However, because the lone pair is closer to the nucleus, it repels the other electrons. This compresses the bond angles slightly, to about 107.5°. This time, the shape of the pyramid, but not a regular means the VSEPR shape of NH ₃ is The shape at right shows the atom downward), plus the "invisible" lone pair of electrons above. H ₂ O has two bonds to hydrogen atoms, and two lone pairs of electrons. The VSEPR shape of the H ₂ O molecule is therefore "bent". Because the lon are closer to the nucleus, they repel a little more strongly than bond electron the bond angle compresses to 104.5°.	Geometry
The "lone pair" of electrons (above the N atom) and the three bonds all repel each other. This gives <u>four</u> electron clouds, just like CH ₄ . However, because the lone pair is closer to the nucleus, it repels the other electrons more strongly than bond electrons. This compresses the bond angles slightly, to about 107.5°. This time, the shape of the pyramid, but not a regular means the VSEPR shape of NH ₃ is The shape at right shows the atom downward), plus the "invisible" lone pair of electrons above. H ₂ O has two bonds to hydrogen atoms, and two lone pairs of electrons. The VSEPR shape of the H ₂ O molecule is therefore "bent". Because the lone angle compresses to 104.5°.	\—H
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The VSEPR shape of the H ₂ O molecule is therefore "bent". Because the lon are closer to the nucleus, they repel a little more strongly than bond electro the bond angle compresses to 104.5°.	riangula This nidal". the N
•	e pairs ons, and
H	

	VSEPR Theory Page: 29	Э2
Big Ideas	Details Unit: Covalent Bonding & Molecular Geomet	ry
	Now, suppose we have a molecule with a double bond, such as CH_2O . The Lewis structure is:	I
	The electrons around the carbon atom are in <u>three</u> clouds, two smaller clouds for the C-H single bonds, and one larger cloud for the C=O double bond. If these bonds got as far apart as possible in 3- dimensional space, they would be the points of a triangle, all in the same plane. This means that CH ₂ O is a "trigonal planar" molecule:	
	Finally, the CO₂ molecule has the following Lewis structure:	į
	It has two large electron clouds for the C=O double bonds. The farthest these clouds can be from each other is 180° apart. This means the molecule forms a straight line, and its VSEPR shape is "linear":	
	The VSEPR shapes in this document are summarized in the table on the following page.	
	Use this space for summary and/or additional notes:	

Big Ideas

Details

Table OF VSEPK Slidpes						
Electron Clouds	Bond Atoms	Lone Pairs	Hybridization	Bond Angle	Picture	VSEPR Shape
4	4	0	sp ³	109.5°		tetrahedral
4	3	1	sp ³	107.5°		trigonal pyramidal
4	2	2	sp ³	104.5°		bent
3	3	0	sp ²	120°		trigonal planar
3	2	1	sp²	118°	, Ö,	bent
2	2	0	sp	180°	0-0-0	linear

		VSEPR Theo	ory	Page: 294		
Big Ideas	Details	Details Unit: Covalent Bonding & Molecular Geometry				
	Homework Problems For each of the following molecules, draw the Lewis structure. Then build a model of the molecule, and use your model to determine the shape of the electron clouds, and the shape of the molecule.					
	Formula	Lewis Structure	# of Electron Clouds around Central Atom	VSEPR shape		
	CHF₃	:F: H—C—F: :F:	4	tetrahedral		
	NI3					
	H ₂ S					
	CO ₂					
	AsH₃					
	COCl ₂					
	PCI ₃					

Unit: Covalent Bonding & Molecular Geometry

Orbital Hybridization

Unit: Covalent Bonding & Molecular Geometry

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Determine the hybridization of the central atom in simple molecules.

Success Criteria:

- VSEPR shapes show the correct number of lone pairs in the correct locations and correct bond angles.
- Hybridization is correct (sp, sp² or sp³).

Tier 2 Vocabulary: hybrid

Language Objectives:

• Explain how electron clouds change shape.

Notes:

orbital: the name for one of the spaces around an atom where electrons are.

<u>hybrid orbital</u>: an orbital whose shape is a hybrid of the shapes of different types of orbitals (such as a cross between an s-orbital and a p-orbital).

It is tempting to think of electrons as well-behaved particles that stay within the rigid boundaries defined by their energy levels. However, electrons are actually tiny charged particles moving randomly at speeds close to the speed of light. Because of their energies and the energies of the nuclei and the other electrons around them, they bounce around within a specific area. If that area is the shared electrons in a covalent bond, the region has a different shape than the electrons of an unbonded atom.

When atoms form covalent bonds, the electrons occupy the space between the two atoms. The space where the bonding electrons are is still called an orbital, even though its shape is now *different* from the shapes of the orbitals in the s, p, d, or f sub-levels of a single atom.

Orbital Hybridization

Page: 296

Big Ideas	Details	, Unit: Covalent Bonding & N	lolecular Geometry
	Recall that molecules v trigonal pyramidal, or l based on a tetrahedral	with four electron clouds (tetrahedral, bent with single bonds, like H2O), are VSEPR shape:	
	The shape of the orbita shape determined by t the following:	Ils surrounding the central atom is the he four electron clouds. It looks like	8.5° -709.5°
	If we wanted to create	four orbitals like this one by reshaping	
	the <i>s</i> and <i>p</i> orbitals of a three <i>p</i> orbitals. We th because it looks like a l	an atom's valent shell, we would need to st perefore call this bonding orbital an sp^3 hyb hybrid made from the one s and three p or	art with one <i>s</i> and rid orbital, bitals.
	Similarly, molecules wi based on the trigonal p	th three electron clouds are lanar VSEPR shape:	
	This hybrid orbital wou p orbitals, and would b	ld come from one s and two e called an sp² hybrid orbital :	120°
	Finally, the hybrid orbinit orbinities of the second second second second second second second second second se	tal from one <i>s</i> and one <i>p</i> orbital is indeed c	alled an sp hybrid

Orbital Hybridization

Big Ideas	Details Unit: Covalent Bonding & Molecular				lar Ge	
	Summary of VSEPR Shapes for Hybrid Orbitals					
		Hybridization	VSEPR Shape(s)	Bond Angles		
			tetrahedral	109.5°		
		sp ³	trigonal pyramidal,	107.5°		
			bent	104.5°		
	sn ²	trigonal planar	120°			
		sp	bent	118°		
		sp	linear	180°		

Orbital Hybridization

Big Ideas Details

Homework Problems

For each of the following molecules, draw the Lewis structure. Then build a model of the molecule, and use your model to determine the shape of the electron clouds, and the shape of the molecule.

Formula	Hybrid- ization	Lewis Structure	# of Electron Clouds around Central Atom	VSEPR shape
CHF₃	sp³	::: : н-с :: ::	4	tetrahedral
SCl ₂				
SiO2				
PH₃				
CH₂O				

		01	SituriySiturzut		1 age: 255
Big Ideas	Details		Unit: Covalent Bonding & Molecular Geom		
	Formula	Hybrid- ization	Lewis Structure	# of Electron Clouds around Central Atom	VSEPR shape
	C ₂ H ₂				
	HCN				

NO₃[−]

 BF_3

Use this space for summary and/or additional notes:

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Orbital Hybridization

Details Introduction: Intermolecular Forces **Unit:** Intermolecular Forces Topics covered in this chapter: Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-2 Use the periodic table model to predict and design simple reactions that result in two main classes of binary compounds, ionic and molecular. Develop an explanation based on given observational data and the electronegativity model about the relative strengths of ionic or covalent bonds. **HS-PS1-3** Cite evidence to relate physical properties of substances at the bulk scale to spatial arrangements, movement, and strength of electrostatic forces among ions, small molecules, or regions of large molecules in the substances. Make arguments to account for how compositional and structural differences in molecules result in different types of intermolecular or intramolecular interactions. HS-PS2-7(MA) Construct a model to explain how ions dissolve in polar solvents (particularly water). Analyze and compare solubility and conductivity data to determine the extent to which different ionic species dissolve.

Unit: Intermolecular Forces

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3

Mastery Objective(s): (Students will be able to...)

- Calculate the electronegativity difference between atoms in a bond.
- Identify polar bonds based on electronegativity differences.

Success Criteria:

• Bonds are correctly identified as polar or non-polar based on electronegativity difference.

Tier 2 Vocabulary: polar

Language Objectives:

• Explain how electrons are distributed unevenly in polar bonds.

Notes:

Details

<u>polar</u>: anything with two sides that are opposite with respect to something. For example, a battery is polar because it has a positive and negative end.



<u>polar bond</u>: a covalent bond that has opposite partial charges on each side (one side partially positive and one side partially negative), because of unequal sharing of electrons.

Dolar Donde

		Polar Bonds	Page: 303
Big Ideas	Details		Unit: Intermolecular Forces
	For example: The bo	nd between H and Cl in the H-C	l molecule is a polar bond.
	The bond is polar be sharing is not equal. electronegativity of o chlorine than with hy	cause hydrogen and chlorine sha Chlorine has an electronegativit only 2.2. This means the electro vdrogen.	are a pair of electrons, but the ty of 3.16, but hydrogen has an ns spend more time with
	One way to show a p the side where the e like this:	olar bond is by using a wedge-sl lectrons spend the most time. T	haped bond, which is wider on The HCl molecule would look
		H⊸Cl	
	The wedge is narrow atom has the electro	er on the H side and wider on th ns more of the time.	ne Cl side because the chlorine
	It is also common to case Greek letter "de would be shown as δ the following examp	label atoms in the structure with elta" (δ) is used to mean "partial" + and a partially negative charge le:	h partial charges. The lower- ". A partially positive charge e would be shown as δ^- , as in
		Å⁺ ⊸ Č	
	In the above example partial negative char	e, hydrogen has a partial positive ge.	e charge, and chlorine has a

Polar Bonds



Polar Bonds

DetailsUnit: Intermolecular ForcesHowever, even covalent bonds can have measurably unequal sharing. it takes
a relatively small difference to create a measurable bond polarity. The lower
boundary is defined by the polarity of a C—H bond, which has no observable
polarity under most conditions. The difference between the Pauling
electronegativity of carbon (2.55) and hydrogen (2.20) is 0.35, so $\Delta \chi = 0.35$ is
chosen to be the maximum electronegativity difference for a nonpolar bond.
This represents about 3 % ionic character.

These values are summarized in the following table:

Electronegativity Difference (Δχ)	% lonic Character	Bond Character
0.35 or less (C—H)	< 3 %	nonpolar covalent
between 0.35 and 1.7	3%-50%	polar covalent
1.7 or more	> 50 %	ionic

Use this space for summary and/or additional notes:

Big Ideas



Use this space for summary and/or additional notes:

Homework Problems

Complete the table. You will need to look up electronegativity values (χ) from in "Table Z. Selected Properties of the Elements" of your Chemistry Reference Tables, which begins on page 516.

Elements	Bond Type	χ1	χ2	Δχ	Bond Character
Pb–S	ionic	2.33	2.58	0.25	nonpolar covalent
Ag–Cl					
Cu–C					
C–N					
C–I					
H–O					
Al–Cl					
K–F					
N–H					
N–O					
C–S					
BaCl					
S-O					
Si–H					

Use this space for summary and/or additional notes:

Big Ideas

Details

Big Ideas

Polar Molecules

Unit: Intermolecular Forces

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3

Mastery Objective(s): (Students will be able to...)

• Draw polarity arrows indicating polarity of a molecule.

Success Criteria:

• Arrow is in the correct direction and points from the δ + atom to the δ - atom.

Tier 2 Vocabulary: polar

Language Objectives:

• Explain how electrons are distributed unevenly in polar molecules.

Notes:

<u>polar bond</u>: a covalent bond that has opposite partial charges on each end (one end partially positive and one end partially negative), because of unequal sharing of electrons.

<u>polar molecule</u>: a molecule that can be oriented so that it has opposite charges on opposite sides.

In order to be polar, a molecule must have both:

- 1. one or more polar bonds
- 2. an "axis of asymmetry," meaning a way to orient the molecule so that there is more partial positive charge on one side (relative to the central atom), and more partial negative charge on the opposite side.





Big Ideas	Details	Unit: Intermolecular Forces
	Dipole Moment	
	The polarity of a molecule can be expressed quantitativel	y as its dipole moment.
	<u>moment</u> : in physics, the degree to which mass is spread (potential) rotation. For example, an object's momer much the object resists forces that would cause it to	out from the center of nt of inertia measures how rotate.
	<u>dipole moment</u> : a measure of how strongly a dipole will r The dipole moment is expressed as the moment of in charges (+q and -q) separated by a distance d.	eact to an external field. ertia caused by a pair of
	The dipole moment (μ) is expressed by the formula:	
	$\mu = qd$	
	The unit for dipole moment is the debye (D).	
	Stronger charges (or partial charges) and/or greater dista will result in a larger dipole moment (and therefore a mo	nce between those charges re polar molecule).

Use this space for summary and/or additional notes:



Unit: Intermolecular Forces

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3, HS-PS2-7(MA)

Mastery Objective(s): (Students will be able to...)

• Rank attractions from strongest to weakest based on the type of intermolecular force.

Success Criteria:

• Attractions are correctly identified and correctly ranked.

Tier 2 Vocabulary: polar

Language Objectives:

• Explain the different types of intermolecular forces and their relative strengths.

Notes:

intramolecular forces: forces within a molecule (chemical bonds)

intermolecular forces ("IMFs"): forces between molecules (solids & liquids). Weaker than intramolecular forces.

Note that both intramolecular forces (chemical bonds) and IMFs form because the process of these atoms or molecules coming together releases energy. If you wanted to separate the atoms/molecules/particles, you would need to add enough energy to make up for the amount of energy that was released.

Forming bonds <u>always</u> releases energy. Breaking bonds <u>always</u> requires energy.

<u>soluble</u>: when the attraction between solvent molecules and solute molecules or ions is strong enough to keep the solute distributed throughout the solvent.

miscible: when two or more liquids are soluble in each other.

Recall the 3 types of compounds:

ionic: compound made of ions (usually metal + nonmetal), which have charges with integer values (±1 or more)

covalent: compound made by sharing of electrons (usually all nonmetals),

metallic: compound made of metal atoms with delocalized electrons

Big Ideas	Details Unit: Intermolecula	ar Forces
	Types of intermolecular forces (IMF), strongest to weakest:	
	The stronger the IMFs, the <u>higher</u> the melting and boiling point of the composite because you have to overcome the IMF in order to separate the molecules g from solid \rightarrow liquid or liquid \rightarrow gas.	ound, oing
	ion-ion: force of attraction between ions. The strength of the force is based Coulomb's Law:	on
	$F = \frac{kq_1q_2}{d^2}$	
	where k is a constant, q_1 and q_2 are the strengths of the two charges, and the distance between them. Bigger charges (larger values of q) mean stu- forces. (<i>E.g.</i> , the attraction between a +2 ion and a -3 ion will produce that's six times as strong as the attraction between a +1 ion and a -1 ion charges are the same, smaller molecules (smaller value of d) have strong forces.	d <i>d</i> is ronger a force .) If ger
	metallic bonds: metal atoms that delocalize their electrons and are held tog the "sea" of electrons surrounding them.	ether by
	dipole-dipole: the force of attraction between two polar molecules (dipoles). Recall
	that the strength of attraction is based on the <u>dipole moment</u> (μ) of the molecule, given by the formula: $\mu = ad$	
	The partial charge (q) is produced by the electronegativity difference (Δ between the two atoms of a polar bond.	X)
	<u>hydrogen bonds</u> : the strongest type of dipole-dipole forces. Occurs in mole that contain hydrogen (χ = 2.20) and an element with an electronegativi than 3.0 (F, O, Cl, or N). The hydrogen bonds that hold water molecules together are what give water its unusual properties:	cules ty larger
	• Water is more dense as a liquid than as a solid.	
	 Water has an unusually high heat capacity (specific heat). 	
	 Water has a relatively high melting and boiling point. (Almost all covalent compounds with a molecular weight as light as 18 amu a gases.) 	are
	 Water exhibits high surface tension and capillary action. 	
	 Water is known as the "universal solvent". 	
	Use this space for summary and/or additional notes:	

Big Ideas	Details		Unit: Int	ermolecular Forces	
	Londor ten (Na for pre mo hav There c <u>ion-dip</u> a d dip <u>induced</u> exp ten	 London dispersion forces (induced dipoles): random movement of electrons causes temporary dipoles to form within molecules, causing very weak attraction. (Named after the chemist Fritz London.) All molecules have these dispersion forces, but they can only be observed when there aren't any stronger forces present. Because dispersion forces are attractions between electrons, molecules with more electrons (generally those with higher molecular mass) have stronger dispersion forces. There can also be IMFs between different types of molecules: ion-dipole attraction: attraction between an ion (full positive/negative charge) and a dipole (partial positive/negative charge). This attraction is stronger than dipole-dipole forces but weaker than ion-ion forces. induced dipole: attraction between a permanent dipole and a compound that experiences only dispersion forces. The permanent dipole induces (creates) a temporary dipole in the other compound 			
			Turns of Common d	Strength	
		Intermolecular Force	Type of Compound	Strength	
		ion-ion	ionic (metal + nonmetal)	strongest	
		metal-metal	metallic (all metals)		
		hydrogen bonds (strong dipole-dipole)	H with F, O, Cl, or N		
		dipole-dipole forces (other than hydrogen bonds)	polar covalent (all nonmetals)	▶ ▶	
		dispersion	nonpolar covalent (all nonmetals)	weakest	
			ditional notae:		

Polar vs. Nonpolar Solvents Because the molecules polar liquids (especially those with hydrogen bonds) at each other, polar liquids will:	tract
Because the molecules polar liquids (especially those with hydrogen bonds) at each other, polar liquids will:	tract
	will
 dissolve other polar liquids (The two liquids are said to be <u>miscible</u>.) 	will
 dissolve ions (from ionic compounds that can separate). E.g., NaCl ions dissolve in H₂O. 	vviii
In general, polar liquids will not dissolve nonpolar liquids or other uncharg molecules.	ed
In general, polar liquids can dissolve ionic compounds that have relatively ion-ion forces (such as ions with $\pm 1 \text{ or } -1$ charges). In general, most polar liquids cannot dissolve most ionic compounds in which the charges of all c ions are ± 2 or higher. (However, note that this is a rule of thumb and the <u>many</u> exceptions!)	small f the e are
Because polar liquids form dipole-dipole bonds, they will squeeze out nonpola liquids. If you mix a nonpolar liquid and a polar liquid (such as oil and water), liquids will form two separate phases. The popular expression to describe this phenomenon is "like dissolves like."	r the

Big Ideas	Details	Unit: Intermolecular Forces		
	Homework Problems			
	For each pair of compounds:Write down the strongest intermolecular force that occurs in each of the two			
	compou	inds.		
	 Circle tr boiling 	ne compound of each pair that would have the higher melting and point (stronger intermolecular forces).		
	Give the	e reason(s) for your choice.		
	1 N-C			
	1. NaCi	HCI		
	2 626			
	2. Cd5			
	3 H-0	N-O		
	5. 1120	1420		
	4. CO	CO ₂		
	5. HBr	HF		
	6. CH ₄	C ₈ H ₁₈		

Introduction: The Mole Unit: The Mole 320 Topics covered in this chapter: 328 Moles 334 Standards addressed in this chapter: Massachusetts Curiculum Frameworks & Science Practices (2016): HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.	Big Ideas	Details Unit: The Mole
Unit: The Mole Topics covered in this chapter: Moles 320 Percent Composition & Empirical Formula 328 Hydrates 334 Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.		Introduction: The Mole
Topics covered in this chapter: 320 Percent Composition & Empirical Formula 328 Hydrates 334 Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.		Unit: The Mole
Moles 320 Percent Composition & Empirical Formula 328 Hydrates 334 Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-P51-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.		Topics covered in this chapter:
Percent Composition & Empirical Formula 328 Hydrates 334 Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.		Moles
Hydrates		Percent Composition & Empirical Formula
Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.		Hydrates
Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.		Standards addressed in this chapter:
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		HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.

Moles

Unit: Moles

Details

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3

Mastery Objective(s): (Students will be able to...)

- Determine the molar mass of a compound.
- Convert between moles, mass, volume, and molecules/atoms.

Success Criteria:

- Conversions are set up properly so undesired units are canceled and desired units appear in the correct place.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: mole, molar

Language Objectives:

• Explain the concept of a quantitative collective noun (such as "dozen") and apply it to the mole concept.

Notes:

<u>mole</u>: (working definition) the amount a compound that is the same number of grams^{*} as the compound's formula or molecular mass in amu.

<u>mole</u>: (formal definition) the amount of matter that contains the same number of objects (atoms, molecules, *etc.*) as the number of atoms in <u>exactly</u> 12 g of ¹²C. When the British Imperial system was more commonly used, this quantity was often called a gram-mole.

<u>Avogadro's constant</u>: 1 mole = 6.022×10^{23} atoms, molecules, *etc.* (*Memorize this number*!)

<u>molar volume</u>: the space occupied by 1 mole of ANY gas. At S.T.P. (0 °C = 273 K) and 1 bar of pressure), the molar volume is 22.7 L. (*Memorize this number!*)

<u>molar mass</u> (m.m.): the mass (in grams) of 1 mole of a substance. For atoms, this is the same number as the atomic mass on the periodic table, but with the unit "grams". For compounds, add up the mass of each atom in the compound.

^{*} Technically, this is only true for moles in the S.I. system, which used to be called "grammoles". In the British Imperial system, a "pound-mole" is the amount of a compound that is the same number of *pounds* as its formula mass. Chemists use the term "ton-mole" the same way other people might use "@#%\$-ton".

	Moles	Page: 321		
Big Ideas	Details	Unit: Moles		
	How Moles are Used in Chemistry			
	Moles are used as a way to make sure you have the desired numbe molecules for a chemical reaction.	r of atoms or		
	For example, consider the reaction:			
	2 Fe + 3 I₂ → 2 FeI₃			
	Suppose you wanted to perform this reaction in a lab with exact an starting materials (for example, let's say it's because you didn't war anything left over). You would need a ratio of 2 atoms of Fe for ever of I ₂ . The problem is that all you have to work with is a jar of iron, a and a balance.	nounts of your ht to have ery 3 molecules a jar of iodine,		
	To solve this problem, we define a mole as being the same number average atomic mass. This means that a mole is always the same n or molecules (Avogadro's constant), so if we start with exactly 2 <u>moles</u> exactly 3 <u>moles</u> of I_2 , we will end up making exactly 2 <u>moles</u> of FeI ₃ .	of grams as the umber of atoms o <u>les</u> of Fe and		
	You can think of a mole the same way you think of a dozen (or a 12 have 3 dozen eggs and 6 dozen slices of toast, you could serve $\frac{1}{12}$	-pack). If you dozen eggs (one		
	egg) and $\frac{2}{12}$ dozen slices of toast (2 slices) to each person. Similar 2 moles of Fe (12.04 × 10 ²³ atoms), and 3 moles of I ₂ (18.06 × 10 ²³ n can react them to make 2 moles of FeI ₃ (12.04 × 10 ²³ molecules).	ly, if you have nolecules), you		
	In the chapter on "Stoichiometry," starting on page 409, we will use in chemical equations to calculate the number of moles of reactant moles of products produced in a reaction.	e the coëfficients s used and the		
	Lise this space for summary and/or additional potes:			

	Mo	oles	Page: 322			
Big Ideas	Details		Unit: Moles			
	Molar Mass					
	The molar mass is the mass (in grams) of one mole of a chemical. Molar mass is used to convert between grams and moles. (Recall the "Conversions (Factor-Label Method)" section on page 88.)					
	 The molar mass of an atom is its (average) atomic mass, but expressed in gram instead of atomic mass units (amu): atom of Fe has a mass of 55.8 amu (from the periodic table), so mole of Fe has a mass of 55.8 g The molar mass of a compound is the sum of the (average) atomic masses of the elements in the compound, but again expressed in grams instead of amu: molecule of I₂ has 2 atoms of I. One atom of I has a mass of 126.9 amu, so molecule of I₂ has a mass of 2 × 126.9 = 253.8 amu. This means that mole of I₂ has a mass of 2 × 126.9 = 253.8 g 					
	Similarly:					
	1 mole of FeI₃ contains 1 mole all of the atoms in the molecu	1 mole of FeI₃ contains 1 mole of Fe atoms and 3 moles of I atoms, so we add up all of the atoms in the molecule:				
	1 mole Fe = 1 × 55.8 = + <u>3 moles I = 3 × 126.9 =</u> 1 mole FeI ₃ =	55.8 g <u>380.7 g</u> 436.5 g				
	Use this space for summary and/o	or additional notes:				

	Moles	Page: 323	
Big Ideas	Details	Unit: Moles	
	Homework Problems		
	Calculate the mass in grams of one mole of each of the following	ng compounds.	
	1. HCl		
	Answer: $36.46 \frac{g}{mol}$		
	2. Fe ₂ O ₃		
	Answer: 36.46 ^{_g}		
	5. 20 0014		
	Answer: $36.46 \frac{g}{mol}$		
	4. $C_6H_{12}O_6$		
	Answer: 36.46^{-g}		
	$5 (PO_{t})_{t}$		
	5. Ca3(1 C4)2		
	Answer: 36.46 ^g mol		

Big Ideas	Details	Unit: Moles
	6. UOCl ₂	
	Answer: 36.46 $\frac{g}{mol}$ 7. NiCl ₂	
	Answer: 36.46 ^g mol 8. (NH ₄) ₂ SO ₄	
	Answer: 36.46 $rac{g}{mol}$ 9. AgNO3	
	Answer: 36.46 ^g 10. CH₃COOH	
	Answer: 36.46 $\frac{g}{mol}$	

Use this space for summary and/or additional notes:
Mo	les

Details Mole Conversions (add up the mass of the formula) 1 mol = ___ grams $1 \text{ mol} = 6.02 \text{ x} 10^{23} \text{ atoms, molecules, etc.}$ $n \text{ mol} = \frac{PV}{PT}$ (1 mol of gas @ S.T.P.* = 22.7 L) These conversions work just like the ones from earlier in the year. Sample Problems: 1. 2.5 mol of NH₃ gas occupies what volume at S.T.P.? $\frac{2.5 \text{ mol}\text{NH}_3}{1} \times \frac{22.7 \text{ L}\text{NH}_3}{1 \text{ mol}\text{NH}_3} = 57 \text{ mol}\text{NH}_3$ 2. What is the mass of 4.1 mol NH_3 gas? The molar mass of 1 mol NH₃ is $(1 \times 14) + (3 \times 1) = 17$ g NH₃. $\frac{4.1 \text{ mol NH}_3}{1} \times \frac{17 \text{ g NH}_3}{1 \text{ mol NH}_3} = 70. \text{ g NH}_3$ 3. How many molecules are there in 0.75 mol of NH₃ gas? $\frac{0.75 \text{ mol}\text{NH}_3}{1} \times \frac{6.02 \times 10^{23} \text{ molecules } \text{NH}_3}{1 \text{ mol}\text{NH}_3} = 4.5 \times 10^{23} \text{ molecules } \text{NH}_3$ 4. What is the volume of 25.5 g of NH₃ gas at S.T.P.? The molar mass of 1 mol $NH_3 = (1 \times 14) + (3 \times 1) = 17 \text{ g } NH_3$. $\frac{25.5 \text{gNHf}_3}{1} \times \frac{1 \text{molNHf}_3}{17 \text{gNHf}_2} \times \frac{22.7 \text{LNH}_3}{1 \text{molNHf}_2} = 34.1 \text{LNH}_3$ Note: this chapter is a good time to start including the chemical formula as part of the units. This will be extremely useful when we study stoichiometry. * S.T.P. = "Standard Temperature and Pressure". Since 1980, the official IUPAC definition of S.T.P. has been 0 °C and 100 kPa. Some texts (and the MA DESE) stubbornly insist on using the old definition of 0 °C and 1 atm. This would make the molar volume of an ideal gas

Use this space for summary and/or additional notes:

22.4 L instead of 22.7 L.

Big Ideas

Dialdaca	Deteile	Moles	Page: 326
Big ideas	Details	Homowork Brobloms	Unit: Moles
	1		
	1.	How many moles are 65.0 grams of zinc?	
	2	Answer: 0.99 mol Zn	
	Ζ.	How many moles are 1250.5 g of lead (II) hitrate $(PD(NO_3)_2)$?	
	2	Answer: 3.775 mol lead (II) nitrate	
	3.	How many moles are 2 500 g of tin (10) chlorate (Sh(ClO ₃) ₄)?	
		Answer: 5.5 mol tin (IV) chlorate	
	4.		
	5	Answer: 0.7357 mol sliver nitrate	toxide?
	5.		
		Answer: 0 570 mol dinitrogon pontovido	
	6.	How many oxygen atoms are there in 380 g of copper (II) phosp	hate?
	_	· · · · · · · · · · · · · · · · · · ·	
		Answer: 1.00 mol conner (II) phosphate	
	7.	How many hydrogen atoms in 454 g of aluminum hydroxide?	
		Answer: 5.82 mol aluminum hydroxide	

Big Ideas	Details	Unit: Mole
	8.	What is the mass (in grams) of 2.35 mol of S_2N_3 ?
	9.	Answer: 249 g S_2N_3 What is the mass (in grams) of 0.25 mol of silver acetate?
	10.	Answer: 42 g silver acetate What is the mass (in grams) of a 2.00 kg bag of table sugar ($C_{12}H_{22}O_{12}$)?
	11.	Answer: 2000 g ☺ How many moles are in 123.5 L of oxygen gas at S.T.P.?
	12.	Answer: 5.44 mol oxygen How many moles are in a 40. gallon drum of chlorine gas at S.T.P.? (1 gal = 3.78 L)
	13.	Answer: 6.7 mol chlorine gas What is the volume (in liters) of 3.5 mol of argon gas at 1.1 atm and 20 °C? (<i>Hint: this is not at S.T.P., so you need to use PV = nRT</i> .)
	14.	Answer: 76.5 L argon gas What is the volume (in liters) of 4.90 \times 10^{25} molecules of N_2 gas at S.T.P.?
		Answer: 1850 L N ₂ gas

Percent Composition & Empirical Formula

Unit: Moles

Details

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3

Mastery Objective(s): (Students will be able to...)

- Determine the percentage of each element in a compound given its chemical formula.
- Determine the empirical and molecular formulas of a compound from percent composition data.

Success Criteria:

- Percentages are calculated correctly.
- Empirical and molecular formulas are calculated correctly.
- Subscripts in empirical and molecular formulas are whole numbers.
- Ratio of subscripts in empirical formulas is in lowest terms.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: mole, composition

Language Objectives:

• Accurately describe the process for converting percentages to moles.

Notes:

percent composition: the percentage by mass of each element in a compound.

molecular formula (chemical formula): a formula that gives the numbers and types of atoms in a molecule.

- <u>empirical formula</u>: a chemical formula with the subscripts reduced to lowest terms. *E.g.*, the empirical formula for C_2H_4 would be CH_2 . The empirical formula for C_8H_{16} would *also* be CH_2 . (You may remember that we <u>always</u> use empirical formulas for ionic compounds.)
- <u>formula mass</u>: the mass in grams represented by a chemical formula. Sometimes called <u>molecular mass</u>, <u>formula weight</u> or <u>molecular weight</u>. (This is the same number as the molar mass, but with units of atomic mass units (amu) instead of grams.)

Percent Composition & Empirical Formula

Details
Determining Percent Composition
To determine the percent by mass of each element in a compound:
1. Determine the atomic mass of the element of interest
2. Determine the formula mass of the entire compound.
3. percent composition = $\frac{\text{atomic mass of element of interest}}{\text{formula mass of entire compound}} \times 100$
Sample Problem:
Q: What is the percentage of carbon in the compound $C_6H_{12}O_6$?
A: Mass of $C_6 = 6 \times 12.01 = 72.06$
Mass of $C_6H_{12}O_6$:
$C_6 = 6 \times 12.01 = 72.06$
$H_{12} = 12 \times 1.008 = 12.096$
$+ O_6 = 6 \times 16.00 = 96.00$
180.156
$\frac{\text{mass of C}_6}{\text{mass of C}_6} = \frac{72.06}{200} = 0.400 \times 100 = 40.0\%$
mass of $C_6 H_{12} O_6 = 180.156$
Determining Empirical and Malagular Formu

Determining Empirical and Molecular Formulas

The lowest-terms ratio of the atoms in a chemical formula is the empirical formula.

The ratio of atoms is the same as the ratio of moles, which means you can find the empirical formula of a compound by determining the ratio of moles, and converting that ratio to whole numbers:

- 1. Find the molar mass (in grams) of each element in the compound.
- 2. Convert the grams to moles for each element.
- 3. Convert the number of moles of each element to whole-number subscripts. The easiest way to do this is by dividing them all by whichever number is the smallest.
- If a subscript is within ±5 % of a whole number after dividing, you can round it off. If a subscript is not within ±5 % of a whole number, multiply <u>all</u> of the subscripts by the smallest number that would cause <u>all</u> of the subscripts to be whole numbers (within ±5 %).

Use this space for summary and/or additional notes:

Big Ideas

Big Ideas	Details U	Jnit: Moles
	Sample Problem:	
	A sample of a chemical compound contains 8.56 g of carbon and 1.44 g of What is the empirical formula of this compound?	hydrogen.
	1. Masses are C: 8.56 g and H: 1.44 g. Write the formula as $C_{8.56g}H_{1.}$	44 g
	2. Convert grams to moles: C: $\frac{8.56 \text{ g}}{1} \times \frac{1 \text{ mol}}{12.011 \text{ g}} = 0.713 \text{ mol}$ U: $\frac{1.44 \text{ g}}{1.44 \text{ g}} = 1 \text{ mol}$ 1.420 mel	
	H: $\frac{1}{1} \times \frac{1.008 \text{g}}{1.008 \text{g}} = 1.429 \text{mol}$	
	The formula for this compound is therefore a whole-number ratio equals $C_{0.713}H_{1.429}$	that
	3. Convert the subscripts to simple whole numbers. The easiest way is to divide them all by the smallest one and see what happens. $C_{\underline{0.713}} \underbrace{H_{1.429}}_{0.713} = C_1 H_{2.004}$	r to do this
	 Round the empirical formula off. CH_{2.004} becomes CH₂. (You can- should—round, as long as you are within ±5 %.) 	–and
	Hints:	
	If the problem gives percentages instead of actual mass, just pretend the percentages are out of 100 g total. <i>E.g.</i> , if you had a compound containing nitrogen, you would use 25.3 g of nitrogen in your calculations.	g 25.3 %
	If you have something like NO _{2.5} , you can't round 2.5 off to 2 or 3. Instead to multiply both subscripts by 2, which gives you N_2O_5 . (This means it's imbe able to recognize decimal equivalents for simple fractions, such as 0.50	d, you need aportant to $0 = \frac{1}{2}$,
	$0.33 = \frac{1}{3}, \ 0.25 = \frac{1}{4}, \ 0.20 = \frac{1}{5}, \ etc.)$	

Percent Composition & Empirical Formula Page 331

	refeelite composition & Empirican ormala	rage. JJI
Big Ideas	Details	Unit: Moles
	Empirical Formula vs. Molecular (actual) Formula	
	If you know the molar mass of the compound, you can use it to get from empirical formula to the molecular formula.	1 the
	For example, suppose you were told that the actual molar mass of the h from the example above is $42.08 \frac{g}{mal}$.	ydrocarbon
	The empirical formula mass (<i>i.e.,</i> the molar mass of the empirical formu $(1 \times 12.011) + (2 \times 1.008) = 14.027.$	la CH ₂) is
	The actual molar mass of 42.08 is 3 times as much, <i>i.e.</i> , $\frac{42.08}{14.027} = 3.00$.	
	This means the molecule contains exactly 3 of the empirical formula uni need to multiply all of the subscripts by 3 to get the molecular formula:	ts, so we
	$CH_2 \times 3 = \boxed{C_3H_6}.$	

Details	5 Unit: Moles
	Homework Problems
1.	A 5.00 g sample of a compound was found to contain 1.93 g carbon, 0.49 g hydrogen and 2.58 g sulfur. What is the empirical formula of the compound?
	Answer: C ₂ H ₈ S
2.	What is the percentage composition of each element in the compound tetrahydrocannabinol (THC), which has the formula $C_{21}H_{30}O_2$?
	Answers: C: 80.2 %; H: 9.6 %; O: 10.2 %
3.	A sample of a compound was found to contain 42.56 g of palladium (Pd) and 0.80 g of hydrogen. If the molar mass of the compound is $2216.8 \frac{g}{mol}$, what is the molecular formula of the compound?
	Answer: Pd ₂ H ₄
4.	Find the empirical formula of a compound that contains 30.45% nitrogen and 69.55% oxygen.
	Answer: NO ₂

Big Ideas

Unit: Moles **Big Ideas** Details 5. Find the percentage of boron in the compound boron triïodide (BI₃). Answer: 2.76 % boron 6. A compound containing only carbon and hydrogen has a molecular mass of 114.26 amu. If one mole of the compound contains 18.17 g of hydrogen, what is its molecular formula? Answer: C₈H₁₈ 7. Find the molecular formula of a compound that contains 56.36 g of oxygen and 43.64 g of phosphorus. The molecular mass of the compound is 283.9 amu. Answer: P₄O₁₀ 8. The compound caffeine has a molecular weight of 194.1926 amu. It contains 49.5% carbon, 5.2% hydrogen, 28.9% nitrogen, and 16.5% oxygen. What is its empirical formula? What is its molecular formula? Answers: empirical: $C_4H_5N_2O$; molecular: $C_8H_{10}N_4O_2$

Hydrates

Unit: Moles

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3 Mastery Objective(s): (Students will be able to...)

• Determine the number of water molecules in a hydrate

Success Criteria:

- Empirical formula is calculated correctly (if necessary).
- Number of water molecules in each formula unit is calculated correctly.
- $\bullet\,$ Formula of hydrate is written correctly, with the empirical first, then a dot, then the number of H_2O molecules.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: hydrate

Language Objectives:

- Explain the concept of a hygroscopic compound.
- Explain the process of determining the amount of water in a hydrate, both numerically and experimentally.

Notes:

hydrate: an ionic solid that has H₂O molecules loosely bound to its crystals.

water of hydration: the water molecules that are bound into a hydrate.

<u>anhydrous</u>: a compound that has had its water of hydration removed, usually by heating.

<u>hygroscopic</u>: a compound that can absorb water from the air. In a humid environment, an anhydrous compound will absorb water until it becomes the hydrate.

Big Ideas	Details Unit: Moles
	Naming of Hydrates
	The name of a hydrate is the name of the compound followed by a number prefix and the word "hydrate".
	The number prefix (the same ones we used for molecular compounds; found in "Table J. Number Prefixes" on page 512 of your Chemistry Reference Tables) indicates the number of H_2O molecules in the hydrate. For example, the compound nickel (II) chloride forms a hydrate that contains six water molecules. Its name is therefore:
	nickel (II) chloride hexahydrate
	Chemical Formula of a Hydrate
	The chemical formula of a hydrate is the chemical formula of the compound followed by a dot and the number of H_2O molecules bound to it. For example, the chemical formula of nickel (II) chloride hexahydrate is:
	$NiCl_2 \cdot 6 H_2O$
	Molar Mass of a Hydrate
	The molar mass of a hydrate includes the mass of the water of hydration. This means a hydrate will have a larger molar mass than the anhydrous compound. For example, the molar mass of NiCl ₂ is 129.60 g. The molar mass of H ₂ O is 18.015 g. The molar mass of 6 H ₂ O molecules is $6 \times 18.015 = 108.09$ g.
	Therefore, the molar mass of NiCl ₂ \cdot 6 H ₂ O is 129.60 + 108.09 = 237.69 g
	Experimentally Determining the Water of Hydration
	You can figure out the formula of a hydrate by weighing it, heating it to remove the water of hydration, and figuring out how many moles of water were removed for every mole of the compound.

Hydrates

Big Ideas	Details	Unit: Moles
	Sample Problem:	
	Q: Sodium sulfate forms a hydrate. We want to find the chemical form hydrate. Suppose you weighed out 32.22 g of the hydrate. After hydrate all of the water, the final mass was 14.20 g.	nula of the eating it to
	A: The formula of anhydrous sodium sulfate is Na ₂ SO ₄ , which has a mo 142.05 g. We have:	olar mass of
	$\frac{14.20 \text{g} \text{Na}_2 \text{SO}_4}{1} \times \frac{1 \text{mol} \text{Na}_2 \text{SO}_4}{142.05 \text{g} \text{Na}_2 \text{SO}_4} = 0.10 \text{mol} \text{Na}_2 \text{SO}_4$	I
	The amount of water removed was 32.22 – 14.20 = 18.02 g.	
	This 18.02 g of H_2O is:	
	$\frac{18.02\text{g}\text{H}_2\text{O}}{1} \times \frac{1\text{mol}\text{H}_2\text{O}}{18.0152\text{g}\text{H}_2\text{O}} = 1.000\text{mol}\text{H}_2\text{O}$	
	Our sample had 1 mole of H_2O and 0.1 mole of Na_2SO_4 . This is 10 ti H_2O as Na_2SO_4 :	mes as much
	$\frac{1 \text{mol}\text{H}_2\text{O}}{0.1 \text{mol}\text{Na}_2\text{SO}_4} = \frac{10 \text{H}_2\text{O}}{1 \text{Na}_2\text{SO}_4}$	
	Therefore, the formula must be:	
	Na ₂ SO ₄ · 10 H ₂ O	

Big Ideas	Hydrates Details	Page: 337 Unit: Moles
	Homework Problems	
	1. What is the chemical formula of iron (III) chloride hexahydrate?	
	2. Give the stock name and molar mass of the compound $CoSO_4$ ·	7 H₂O?
	 If 10.0 g of Na₂CrO₄ · 4 H₂O is heated to constant mass (<i>i.e.</i>, until water of hydration is removed), what will the final mass be? 	all of the
	a. Find the molar mass of $Na_2CrO_4 \cdot 4H_2O$.	
	b. Find the number of moles of Na2CrO4 \cdot 4 H2O in 10.0 g.	
	c. Find the molar mass of anhydrous Na ₂ CrO ₄ .	
	 d. Convert the number of moles (<i>which you found in part b</i>) to using the molar mass of the anhydrous compound (<i>which y part c</i>). 	o grams, You found in

	Hydrates	Page: 338
Details		Unit: Moles
 14.70 g of a hydrate a mass of 11.10 g at determine the cher 	e of CaCl ₂ is heated to dryness. fter evaporating the H ₂ O. Use nical formula of the hydrate.	The anhydrous sample has the following steps to
a. Find the moles	s of anhydrous compound left	at the end.
b. Find the moles water evapora	s of water evaporated. (You'll ited, and then convert to mole	need to find the grams of es.)

d. Write the formula for the hydrate (*using the number of water malecules* that you found in part c)?

Use this space for summary and/or additional notes:

Big Ideas

		1 4861 999
Big Ideas	Details	Unit: Solutions
	Introduction: Solutions	
	Unit: Solutions	
	Topics covered in this chapter:	
	Solutions & Dissolution	
	Solubility	
	Concentration (Molarity)	354
	Colligative Properties	
	Standards addressed in this chapter:	
	Massachusetts Curriculum Frameworks & Science Practices (2016):
	HS-PS1-11(MA) Design strategies to identify and separate the mixture based on relevant chemical and physical properties of the strategies of the strategie	he components of a ties.
	HS-PS2-7(MA) Construct a model to explain how ions dissol (particularly water). Analyze and compare solubility and to determine the extent to which different ionic species	ve in polar solvents I conductivity data dissolve.

Unit: Solutions

Details

MA Curriculum Frameworks (2016): HS-PS2-7(MA)

Mastery Objective(s): (Students will be able to...)

- Describe how a solution forms.
- Explain the effect of temperature changes on solubility.

Success Criteria:

- Descriptions account for solvent-solute interactions.
- Descriptions account for intermolecular forces.
- Explanations of the effect of temperature are consistent with solubility curves.

Tier 2 Vocabulary: solution

Language Objectives:

• Explain how solutes dissolve in solvents.

Notes:

<u>solute</u>: a substance that is broken down and dissolved into another substance. Solutes can be solids, liquids, or gases.

solvent: a substance that contains a solute. Solvents can be solids or liquids.

solution: a mixture that consists of a solute dissolved in a solvent.

dissolution or solvation: the process of a solute dissolving in a solvent.

<u>solubility</u>: the amount of a solute that can dissolve in a solvent. Often expressed in $\frac{\text{mol}}{1}$ or $\frac{g}{1}$.

soluble: when a solute can dissolve in a solvent.

<u>insoluble</u>: when a solute cannot dissolve in a solvent. Common threshold values are that solutes with solubilities of less than $1\frac{g}{L}$ or less than $0.01\frac{mol}{L}$ in a given solvent are considered insoluble.

miscible: when two liquids can dissolve in (mix freely with) each other

<u>dissociation</u>: when ions split apart in a solution. *E.g.*, when NaCl dissolves, the Na⁺ and Cl⁻ ions separate and dissolve separately.

Solutions & Dissolution

		1 age: 54.
Big Ideas	Details	Unit: Solution
	electrolyte: a solution that conducts electricity. Electrolytes are get when ionic compounds (salts) dissociate and dissolve, and the id electrons (electricity) through the solution.	nerally made ons conduct
	saturated solution: a solution that holds as much solute as the solve dissolving at a given temperature.	ent is capable of
	unsaturated solution: a solution that contains less solute than is can dissolving in a solvent.	pable of
	supersaturated solution: a solution that temporarily contains more capable of remaining dissolved in a solvent. Supersaturated sol unstable.	solute than is utions are
	A solution forms when solute molecules are dissolved in solvent mo process involves the following steps:	lecules. This
	1. Solvent molecules are attracted to the surface of the solute	
	 Intermolecular bonds (<i>e.g.,</i> ion-dipole bonds, hydrogen bon between solvent and solute particles pull the solute particle molecules, <i>etc.</i>) apart and into the solvent. 	ds, <i>etc</i> .) es (ions,

Use this space for summary and/or additional notes:

Big Ideas	Details Unit: Solutions
	Enthalpy (Heat) of Solution
	If a solute dissolves in a solvent, it is <u>always</u> the case that more energy had to be released when the solvent-solute intermolecular bonds are formed than it took to pull the solute particles apart. This means that the combined intermolecular forces between the solvent and solute particles are stronger than the intermolecular forces that had held the particles together in the solute.
	If a solute does not dissolve, this means it would have taken more energy to pull the solute particles apart than the amount that would have been released by forming the solvent-solute intermolecular bonds. This means that the combined intermolecular forces between the solute particles are stronger than the combined intermolecular forces between solvent and solute particles.
	This energy can exist in two forms: enthalpy (heat) and entropy (how much the energy is spread out among the particles). Enthalpy and entropy are discussed in more detail in the chapter on "Thermochemistry (Heat)," starting on page 433.
	If the solution gets hotter as the solute dissolves, this means energy was <i>released</i> in the form of enthalpy (heat).
	If the solution gets colder as the solute dissolves, this means heat energy was <i>absorbed</i> . However, it still must be true that energy had to be released when the solute dissolved. (Otherwise it would not have done so.) This means that entropy must have increased, and that more energy was released in the form of entropy than was absorbed in the form of enthalpy (heat).

Solutions & Dissolution

Big Ideas Details **Unit: Solutions** For example, if you mix a strong acid with sodium hydroxide (a strong base), the solution gets very hot. (In fact, it can get hot enough to boil!) However, if you mix a strong acid with sodium carbonate ("soda ash") or sodium hydrogen carbonate (baking soda), the solution gets cold, because it releases CO₂ gas. As the gas is released, its heat energy spreads out into the surroundings (the room), which is a large increase in entropy. This increase in entropy releases so much energy that it takes thermal energy (heat) away from the solution, cooling it off. This is why baking soda is a good choice for neutralizing strong acids, whereas sodium hydroxide would be a poor choice. HCI + NaOH HCI + NaHCO₃

Polar vs. Non-Polar Solvents

Whether a solute will dissolve in a solvent depends on the intermolecular forces between both the solvent and solute molecules. In both cases, the governing factor is the greater strength of ion-ion and dipole-dipole interactions as compared with London dispersion forces.

Polar Solvent

Details

Big Ideas

<u>polar or ionic solute</u>: polar or ionic solute particles are attracted to the positive and negative poles of the solvent molecules, which results in the solute dissolving.

<u>non-polar solute</u>: non-polar solute particles are not attracted to the solvent molecules. However, the solvent molecules are attracted to each other, and they exclude the solute.

Non-Polar Solvent

<u>polar or ionic solute</u>: polar or ionic solute particles are attracted to each other, but are not attracted to the solvent molecules, so they exclude the solvent and do not dissolve. (They form a precipitate, which means the solute falls (precipitates) to the bottom of the container.)

<u>non-polar solute</u>: neither the solute particles nor solvent molecules are strongly attracted to each other. (Both exhibit only London dispersion forces.) Because neither excludes the other, they spread out and intermingle freely.

A simple one-sentence statement of the above is "Like dissolves like."

This statement applies to liquids as well as solids. Polar liquids are miscible with each other; non-polar liquids are miscible with each other; however, non-polar liquids are not miscible with polar liquids. This is why "oil and water do not mix."

Solutions & Dissolution

Page: 346
Unit: Solutions

Big Ideas	Details Unit: Solution
	Homework Problems
	For each solute given, indicate whether water (H_2O) or cyclohexane (a nonpolar molecule) would be a better solvent.
	1. KNO ₃
	2. paraffin (long-chain hydrocarbons, such as $C_{20}H_{42}$ or $C_{40}H_{82}$)
	3. ethyl alcohol (CH ₃ –CH ₂ –OH)
	5. mineral oil
	6. ammonia (NH₃)
	7. gasoline (short-chain hydrocarbons such as octane, C_8H_{18})

Solubility

Unit: Solutions

Details

MA Curriculum Frameworks (2016): HS-PS2-7(MA)

Mastery Objective(s): (Students will be able to...)

- Use solubility tables/rules to predict whether a solute will dissolve in water.
- Determine the amount of a solute that can dissolve from a solubility curve.

Success Criteria:

- Predictions about dissolution in water are correct.
- Amounts of solute that can dissolve are determined correctly.

Tier 2 Vocabulary: solution, curve

Language Objectives:

• Explain how solutes dissolve in solvents.

Notes:

In class, you saw a demonstration of the reaction between sodium carbonate (Na_2CO_3) and calcium chloride $(CaCl_2)$:

 $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow NaCl(aq) + CaCO_3(ppt)$ (1)

When the solutions were mixed, the calcium carbonate that was formed immediately precipitated (formed an insoluble solid). Note that once the calcium carbonate is formed, it doesn't redissolve. *I.e.*, reaction (1) happens, but the reverse reaction (2), doesn't:

 $CaCO_{3}(s) + NaCI(aq) \xrightarrow{} CaCI_{2}(aq) + Na_{2}CO_{3}(aq)$ (2)

This is because of the way ionic compounds behave when they are dissolved in water.

If an ionic compound dissolves in water, it <u>dissociates</u> (splits) into its ions. In a chemical equation, we write "(aq)" (meaning "aqueous") after an ionic compound to show that it is dissolved, and is floating around in the solution as separate positive and negative ions.

For example, $CaCl_2$ splits into one Ca^{2+} ion and two Cl^- ions. The Ca^{2+} ions are attracted to the negative part of the H_2O molecule (the oxygen atoms), and Cl^- ions are attracted to the positive parts (the hydrogen atoms).

Solubility

Details **Big Ideas Unit: Solutions** The combined attraction between the ions and the water molecules is stronger than the attraction between the Ca²⁺ ion and the Cl⁻ ion. The stronger attraction wins, which means the CaCl₂ dissolves: $CI^{-} H^{+} H^{+} CI^{-} CA^{2+}$ CaCO₃, on the other hand, does not dissociate. This must mean that the attraction between the Ca^{2+} ion and the CO_3^{2-} ion is stronger than the combined attraction between the ions and the water molecules. The stronger attraction wins, which means the CaCO₃ precipitates. Note that if you mix the reactants and all of the ions remain in solution, nothing changes. This means a chemical reaction did not occur. In other words, a chemical reaction in an aqueous solution happens only if one of the products forms its own distinct phase—either a precipitate, a gas, or a separate liquid phase. Use this space for summary and/or additional notes:

Sol	lubi	lity
-----	------	------

Big Ideas	Details Unit: Solutions
	Solubility Rules
	Solubility rules are rules of thumb that describe which compounds are likely to be soluble in water, and which are not.
	Recall that the strength of ion-ion intermolecular forces is given by Coulomb's Law: $F = \frac{kq_1q_2}{d^2}$
	<i>I.e.,</i> the attraction is proportional to the absolute value of the product of the charges ($ q_1q_2 $ — multiply the charges, and then change the sign so that the result is a positive number) and inversely proportional to the square of the distance between the ions.
	It is usually (but not always) true that for the solute:
	 if q₁q₂ ≥ 4, then the <i>ions'</i> attraction to each other is usually stronger, and the compound usually precipitates.
	• if $ q_1q_2 < 4$, then the solvent's attraction to the ions is usually stronger, and the compound usually dissolves.
	Note that there are several exceptions to both of these rules. Two examples are:
	 hydroxides (OH⁻) and fluorides (F⁻) tend to form precipitates with +2 ions because they are very small ions, so the force of intermolecular attraction (F) is stronger because d² is smaller.
	 cations (positive ions) of atoms with electronegativities significantly greater than 1 (such as Cu⁺¹, Ag⁺¹, and Pb⁺²) have a stronger attraction for negative ions, and form precipitates with halogens (Cl⁻, Br⁻, and I⁻).
	Use this space for summary and/or additional notes:

Big Ideas

Details

The following is a detailed set of solubility rules:

Ions That Form SOLUBLE Compounds	EXCEPT with	Ions That Form INSOLUBLE Compounds	EXCEPT with	
Group 1 ions (Li ⁺ , Na ⁺ , <i>etc.</i>)		carbonate (CO ₃ ^{2–})		
ammonium (NH_4^+)		chromate (CrO ₄ ^{2–})	Crown 1 ions NUL +	
nitrate (NO $_3^-$)		phosphate (PO ₄ ^{3–})	Group 1 Ions, NH4	
hydrogen carbonate (HCO₃ [–])		sulfite (SO ₃ ^{2–})		
chlorate (ClO₃ [–])		$cultido (S^{2-})$	Group 1 ions, NH4 ⁺	
perchlorate (ClO ₄ ⁻)		sunde (S)	Group 2 ions	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	Ag⁺	hydroxide (OH⁻)	Group I ions,	
halides (Cl⁻, Br⁻, I⁻)	Ag ⁺ , Cu ⁺ , Pb ²⁺ , Hg ₂ ²⁺	oxide (O ^{2–})	NH4 ⁺ , Ba ²⁺ , Sr ²⁺ , Tl ⁺	
sulfates (SO ₄ ^{2–})	Ca ²⁺ , Sr ²⁺ , Ba ²⁺ , Ag ⁺ , Pb ²⁺			



Solubility

Big Ideas	Details	Unit: Solutions
		Homework Problems
	For the Solubili Tables.	se problems, you will need to use the solubility curves in "Figure I. ties of Selected Compounds" on page 511 of your Chemistry Reference
	1.	How much ammonium chloride could you dissolve in 100 g of water at 70 $^\circ\text{C}?$
		Answer: about 61 g NH₄Cl
	2.	How much HCl could you dissolve in 25 g of water at 45 °C?
		Answer: 15 g HCl
	3.	If you made a saturated solution of ammonia in 40. g of water at 50. °C, how many grams of ammonia would it contain?
		Answer: 12 g NH ₃
	4.	You want to dissolve 0.75 mol of KCl (F.W. = $74.55 \frac{g}{mol}$) in 150. mL of water. What is the minimum temperature to which you would have to heat the water to dissolve all of the KCl?
		Answer: 34 °C

		Solubility	Page: 353
Big Ideas	Details		Unit: Solutions
	5.	You have a solution that contains 43 g of an unknown comp in 100. g of H_2O at a temperature of 55 °C. The unknown co be either KCl, Na_2SO_4 , KNO_3 , or $NaNO_3$. Describe how you c series of heating or cooling experiments and use a solubility the solute in the unknown solution.	ound dissolved mpound could ould perform a chart to identify
	6.	If you had 95 g of a saturated solution of sodium nitrate at r temperature (25 °C) and you cooled it to 10. °C, how much p form?	oom precipitate would
		(Note: the 95 g of solution includes both the NaNO3 and the	r water.)
		Answer: 6 g	

Unit: Solutions

MA Curriculum Frameworks (2016): HS-PS2-7(MA)

Mastery Objective(s): (Students will be able to ...)

- Calculate the concentration of a solution in $\frac{mol}{L}$.
- Calculate the final concentration of a solution after dilution.

Success Criteria:

- Solutions have the correct quantities substituted for the correct variables.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: concentration, molar

Language Objectives:

• Explain how concentration is calculated.

Notes:

<u>concentration</u>: how much of something (solute) is dissolved in something else (solvent).

	moles of solute	mol
molarity (M): a unit of concentration equal to		or — .
	L of solution	L

<u>dilution</u>: the process of decreasing the concentration of a substance by adding more solvent.

dilute: a solution that has a low concentration of solute dissolved in it.

There are three common types of problems involving molarity:

- 1. Find the molarity of a solution containing ____ grams/moles of solute with a volume of ____ L.
- "How many moles/grams of a chemical would you need to make ____ L of a ____ M solution?" OR "What volume would you add to ____ moles/grams of a chemical to make a ____ M solution?"
- What volume of ____ M solution would you add to water to make ____ L of a ____ M solution.

Big Ideas	Det	tails Unit: Sc	olutions
	1.	Determining Concentration	
		To calculate the molarity of a solution:	
		1. find the moles of solute	
		2. find the liters of solution	
		3. divide the moles by the liters	
		For example: Determine the molarity of a solution made by dissolving 0.2 of CuSO₄ in enough water to make a total volume of 500 mL (0.5 L) of solu	25 mol ution.
		$\frac{0.25 \text{ mol } \text{CuSO}_4}{0.5 \text{ L solution}} = 0.5 \frac{\text{mol } \text{CuSO}_4}{\text{L}} = 0.5 \text{ M } \text{CuSO}_4$	
		(pronounced "0.5 molar copper sulfate").	
	2.	Determining the mass of solute or the volume of water needed	
		To solve these problems, use the molarity as a conversion factor, rewritin	g M as
		$\frac{\text{mol}}{\text{L}}$ (for example, 1.75 M would be $\frac{1.75 \text{ mol}}{1 \text{L}}$, or 1.75 mol = 1 L).	
		For example:	
		How many moles of AgNO $_3$ would you need to dissolve in water to make 100. mL (0.100 L) of an 0.50 M solution?	
		$\frac{0.100 \text{L}}{1} \times \frac{0.50 \text{mol}}{1 \text{L}} = 0.050 \text{mol}$	
		Note that if the question had asked for <u>grams</u> of AgNO ₃ , you would then r convert moles of AgNO ₃ to grams.	need to

Big Ideas	De	tails	Unit: Solutions
	3.	Dilution problems	
		The idea is that the moles of solute before dilution must equal the solute after dilution. Because molarity times volume equals mol	ne moles of es
		$\left(\frac{\text{mol}}{L} \times L = \text{mol}\right)$, this means $M_1V_1 = M_2V_2$, where M_1 and M_2 are	e the molarities
		before and after dilution, respectively, and V_1 and V_2 are the volu and after dilution.	umes before
		For example:	
		How much 0.50 M HCl would you need to add to water to make 0.10 M solution?	2.0 L of an
		$M_1V_1 = M_2V_2$	
		<i>M</i> ₁ = 0.50 M	
		$V_1 = V_1$	
		$M_2 = 0.10 \text{ M}$	
		$V_2 = 2.0 L$	
		$(0.50)V_1 = (0.10)(2.0)$	
		$V_1 = \frac{(0.10)(2.0)}{0.50} = 0.40 \mathrm{L}$	

Use this space for summary and/or additional notes:

_

	Concentration (Molarity)	Page: 357				
Big Ideas	Details	Unit: Solutions				
	Homework Problems					
	1. What is the molarity of a solution that contains 25.2 g of KNO	3 (F.W. =				
	101.1 ^g _{mol}) dissolved in enough water to make a total volume of solution?	of 200. mL of				
	Answer: 1.25 M					
	2. What is the molarity of a solution that contains 22.5 g of NaI (1 149.98 ^g / _{mol}) dissolved in enough water to make a total volume solution?	F.W. = of 500. mL of				
	Answer: 0.300 M					

Big Ideas	Details Unit: Solutions
	3. How many grams of NaOH (F.W. = 40.00 ^g / _{mol}) would you dissolve in water to make 1.0 L of a 2.0 M solution?
	Answer: 80. g NaOH
	4. How many grams of KCI (F.W. = 74.55 ^g / _{mol}) would you dissolve in water to make 250. mL of 0.100 M solution?
	Answer: 1.86 g KCl
	5. How many mL of 12.0 M HCl would you add to water to make 500. mL of a 1.00 M solution?
	Answer: 42.0 mL HCl

Big Ideas	Details	Unit: Solutions
	6. If you put two teaspoons (8.0 g) of sucrose (C ₁₂ H ₂₂ O ₁₁) into 30 what is the concentration of sugar in the resulting solution?	0. mL of coffee,
	Answer: 0.078 M	

Big Ideas

Colligative Properties

Unit: Solutions

MA Curriculum Frameworks (2016): HS-PS2-7(MA)

Mastery Objective(s): (Students will be able to...)

- Calculate boiling point elevation, freezing point depression, vapor pressure lowering and osmotic pressure.
- Calculate the molar mass of a solute, based on the grams of a solute added and its effect on the freezing or boiling point of water.

Success Criteria:

- Solutions use the equation appropriate for the information given.
- Solutions have the correct quantities substituted for the correct variables.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: depression, elevation

Language Objectives:

• Explain why solutes cause changes in freezing and boiling point.

Notes:

<u>colligative properties</u>: properties of a solution that depend on the physical number of particles dissolved, but not on the chemical properties of those particles.

Solutes can affect the physical properties of a solution by "getting in the way" of the solvent molecules.

 $\underline{molality}$ (m): the concentration of a solution measured in grams of solute per 1000 g of solvent.

Notice that the molality depends only on the masses of the solute and solvent, not on the volume.

<u>van't Hoff factor</u> (*i*): the number of particles of solute that you get when the solute dissolves. For example, when you dissolve sodium phosphate (Na₃PO₄) in water, it breaks up into three Na⁺ ions and one PO₄³⁻ ion, for a total of four particles. This means the van't Hoff factor for Na₃PO₄ is 4.

Note that the van't Hoff factor (i) is a measured quantity. If an ionic compound dissociates completely, the value of i can be approximated from the chemical formula. However, if a compound is a weak electrolyte (dissolves only partially), the value of i must be measured empirically.
Colligative Properties Details **Freezing Point Depression** When a solute is added to a solvent, the solvent particles must "push" the solute particles out of the way in order to form a solid, which requires energy. This means that in order to make the solution freeze, the temperature must be lower, in order to increase the amount of energy *given off* when the solution forms a solid. This is why we put salt on ice in winter—the salt particles get in the way of the water freezing, which means the temperature has to be lower in order for the salt water to freeze. As long as the temperature is above this new freezing point, the solution stays liquid (*i.e.*, the ice melts). $\Delta T_f = imK_f$ where: i = van't Hoff factor m = molality of the solute $\left(\frac{\text{mol solute}}{\text{kg solvent}}\right)$ K_f = freezing point depression constant. For H₂O, $K_f = 1.86 \frac{°C}{m}$ (degrees Celsius per molal) Sample problem: What is the freezing point of 25 g Na_2SO_4 dissolved in 500 g of H_2O ? $\Delta T_f = imK_f$ i = 3 (because Na₂SO₄ \rightarrow 2 Na⁺ + SO₄²⁻, which is a total of 3 ions)

 $m = \frac{25 \,\mathrm{g} \,\mathrm{Na_2 SO_4}}{500 \,\mathrm{g} \,\mathrm{H_2 O}} \times \frac{1 \,\mathrm{mol} \,\mathrm{Na_2 SO_4}}{142.05 \,\mathrm{g} \,\mathrm{Na_2 SO_4}} \times \frac{1000 \,\mathrm{g} \,\mathrm{H_2 O}}{1 \,\mathrm{kg} \,\mathrm{H_2 O}} = 0.352 \,\mathrm{m}$ $K_f = 1.86 \frac{C}{m}$ $\Delta T_f = imK_f$ $\Delta T_f = (3) (0.352 \text{ m}) (1.86 \frac{C}{m})$ $\Delta T_f = 1.96 \,^{\circ}{\rm C}$ The normal boiling point of H₂O is 0 °C, and we just calculated that the freezing point is *lowered* by 1.96 °C. Therefore, the freezing point of the solution is: $T_f = -1.96 \ ^{\circ}\text{C}$

Use this space for summary and/or additional notes:

Big Ideas

	compative rioperties	1 age: 502
Big Ideas	Details	Unit: Solutions
	Boiling Point Elevation	
	Solute particles attract solvent molecules as they boil and attempt gas. The solution needs extra energy (higher temperature) in order this extra attraction. This is why solutions—liquids with solutes di boil at higher temperatures.	t to escape as a er to overcome ssolved in them—
	$\Delta T_b = im K_b$	
	where:	
	i = van't Hoff factor (# solute particles from each molecule, so the "dissociation factor")	ometimes called
	m = molality of the solute $\left(\frac{\text{mol solute}}{\text{kg solvent}}\right)$	
	K_b = boiling point elevation constant. For H ₂ O, $K_b = 0.52 \frac{°C}{m}$ (oper molal)	degrees Celsius
	Calculations involving boiling point elevation are done exactly the calculations involving freezing point depression.	same way as
	Sample problem:	
	Q: It is often said that salt should be added to boiling water when because the salt will elevate the boiling point of the water, can cook faster. How much would one teaspoon (4 g) of salt raise of 4 quarts (about 4 kg) of water?	n cooking pasta using the pasta to the boiling point
	A: 4 g of NaCl is approximately 0.068 mol.	
	The molal concentration of salt in the water is therefore $\frac{0.068 \text{ mol NaCl}}{4.0 \text{ kg H}_2 \text{O}} = 0.017 \text{ m}.$	
	NaCl dissociates into to ions, so $i = 2$. $K_b = 0.52 \frac{C}{m}$. Therefore	:
	$\Delta T_{b} = imK_{b}$	
	$\Delta T_b = (2)(0.017)(0.52) = 0.018^{\circ}\text{C}$	
	The salt in the water would increase the boiling point from 10 We can therefore discount the possibility that boiling point ele significant contribution to how quickly the pasta cooks.	0 °C to 100.018 °C. evation makes any
	L Use this space for summary and/or additional notes:	

Raoult's Law (Vapor Pressure Lowering)

Solute particles attract solvent molecules. This attraction is strong enough to prevent some of those solvent molecules from escaping into the vapor phase.

Vapor pressure is the number of molecules of liquid that can escape into the gas phase at a given temperature, expressed as a pressure. Therefore, the presence of solute particles lowers the vapor pressure of the solvent.

Specifically, Raoult's Law states that the partial pressure of *vapor* "*i*" (P_i) equals the vapor pressure of (pure) "*i*" ($P_{v,i}^{\circ}$) times the mole fraction of *liquid* "*i*" (χ_i) in the mixture:

 $P_i = P_{v,i}^\circ \chi_i$

Sample problem:

A sealed chamber contains a solution of glucose dissolved in water at 22 °C. The vapor pressure of pure water at 22 °C is 2.6 kPa. If the mole fraction of glucose is 0.10, what is the partial pressure of water in the air space above the solution?

Answer:

Details

Big Ideas

If the mole fraction of glucose in the solution is 0.10, the mole fraction of water in the solution must be 1 - 0.10 = 0.90. Therefore, the partial pressure of water is:

$$P_{H_2O} = P_{v,H_2O}^{\circ} \chi_{H_2O}$$
$$P_{H_2O} = (2.6 \text{ kPa})(0.90) = 2.3 \text{ kPa}$$

Big Ideas	Details	Unit: Solutions	
	Osmotic Pressure		
	Diffusion is the natural flow of molecules from a region of higher concentration to a region of lower concentration.		
	Recall from biology that osmosis is a form of diffusion in which solvent molecules are able to travel across a semi-permeable membrane (such as a cell membrane), but solute molecules cannot pass through. Therefore, the higher the concentration of solute molecules on one side of the membrane, the more strongly those solute molecules attract solvent molecules from the other side. The force of this attraction can be measured as a pressure.		
	H_2O H_2O Salt Water Lower $[H_2O]$ H_2O Higher $[H_2O]$		
	Semi-Permeable Membrane		
	This is why your skin wrinkles when it gets wet—the solutes inside you attract the pure water from outside your skin. As the water flows in membrane, it enlarges your skin cells. As your skin gets larger, the su larger, which we see as wrinkles. As your skin dries, the water escape shrink, and the wrinkles disappear.	our skin cells through the cell ırface area gets es, the cells	

Big Ideas	Details	Unit: Solutions
	<u>osmotic pressure</u> (π): the observed pressure difference across a se membrane because of differences in solute concentration. (Ye to think of the Greek letter π as a variable. Chemists are weird	mi-permeable s, it is awkward)
	Because osmotic pressure is a pressure, and because we are as solute molecules otherwise obey kinetic-molecular theory (<i>i.e.</i> , freely, more or less like gas molecules), we can apply the ideal that we need to include the van't Hoff factor because <i>each</i> of t solution created by dissolving a compound contributes separat osmotic pressure. This gives us the following formula:	suming the they move gas law. Note ne ions in ely to the
	$\pi V = inRT$ where:	
	π = osmotic pressure (additional pressure due to osmosis) V = volume of solution i = van't Hoff factor n = moles of solute R = gas constant T = temperature (Kelvin)	
	Because molarity equals the moles of solute (<i>n</i>) divided by the solution (<i>V</i>), the above equation can be simplified to:	volume of
	$\pi = iMRT$	
	where <i>M</i> = molarity of the solute, and everything else is as abo	ve.

Big Ideas D	etails	0	•	Unit: Solutions
		Homewo	ork Problems	
	 If 45 grams of would the me K_b(H₂O) = 0.52 	sodium chloride v Iting and boiling p $\frac{c}{m}$ and $K_f(H_2O) = \frac{c}{m}$	vere added to 500. oints be of the resu 1.86 ^{°C} _m .	grams of water, what Ilting solution?
	Answer: M.P.	= −5.73 °C B.P. :	= 101.6 °C	
	2. What is the vavor vapor pressur	apor pressure of th e of pure water at	ie solution in probl 250 °C is 3.17 kPa.	em #1 at 250 °C? The
	Answer: 3.08	kPa		
	3. If the solution on one side of were placed of pressure be a	in problem #1 (wl f a semipermeable on the other side o t 27 °C?	hich has a density o membrane, and a f the membrane, w	of 1.056 $rac{g}{mL}$) were placed 1.00 M solution of NaCl hat would the osmotic
	Answer: 12.1	atm		

D's Like		
Big Ideas	Details Unit: Solut	lons
	4. Which solution will have a higher boiling point: a solution containing 105 sucrose (C ₁₂ H ₂₂ O ₁₁) in 500. g of water, or a solution containing 35 g of NaC 500. g of water?	g of Cl in
	Answer: for the sucrose, $T_b = 100.32$ °C for the NaCl, $T_b = 101.40$ °C	
	 0.546 g of a compound with a van't Hoff factor of 1 was dissolved in 15.0 benzene. The freezing point of the solution was found to be 0.240 °C low 	g of er
	than the freezing point of pure benzene. If K_f for benzene is $K_f = 5.12 \frac{\circ C}{m}$,	,
	Answer: 776 ^g mol	

Introduction: Chemical Reactions

Unit: Chemical Reactions

Details

Big Ideas

Topics covered in this chapter:

Chemical Equations	.370
Types of Chemical Reactions	.374
Predicting the Products of Chemical Reactions	.379
Activity (Reactivity) Series	.384
Balancing Chemical Equations	.387
Net Ionic Equations	.395

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

- **HS-PS1-2** Use the periodic table model to predict and design simple reactions that result in two main classes of binary compounds, ionic and molecular. Develop an explanation based on given observational data and the electronegativity model about the relative strengths of ionic or covalent bonds.
- **HS-PS1-7** Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.

Unit: Chemical Reactions

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to ...)

• Read, write, and interpret chemical equations.

Success Criteria:

- Equations have reactants and products on the correct sides of the arrow.
- Physical states, heat, solvents, catalysts, *etc.* are present when appropriate.

Tier 2 Vocabulary: equation

Language Objectives:

• Define the symbols used in chemical equations.

Notes:

chemical equation: a set of symbols that describe a chemical reaction. For example: $2 H_2 (g) + O_2 (g) \xrightarrow{\Lambda} 2 H_2 O(\ell) + heat$

<u>reactants</u>: the starting materials; chemicals (and things like energy) that react. In a chemical equation, the reactants are before the arrow (on the left). In the above equation, the reactants are $H_2(g)$ and $O_2(g)$.

<u>products</u>: chemicals (and other things like energy) that are produced. In a chemical equation, the products are after the arrow (on the right). In the above equation, the products are $H_2O(\ell)$ and heat.

Big Ideas	Details	Unit: Chemical Reactions
	state of matter: t physical state following tabl	he symbols in parentheses after a compound indicates the of that compound. Some of the common ones are listed in the e:
	St	ates of Matter Used in Chemical Equations
	Symbol	Meaning
	(s)	solid
	(ℓ)	liquid (A script "L" is often used to avoid confusion between the letter "I" and the number "1".)
	(g)	gas or vapor
	(cd)	condensed phase (<i>i.e.</i> , either solid or liquid)
	(fl)	fluid phase (<i>i.e.,</i> either liquid or gas)
	(cr)	crystalline (solid is in the form of crystals)
	(lc)	liquid crystal
	(vit)	vitreous (glass-like)
	(ads)	adsorbed onto a substrate
	(sln)	solution
	(aq)	aqueous solution (dissolved in water)
	(am)	amorphous solid
	(ppt)	precipitate (solid) formed by the reaction
	reaction condition make the read and/or below • Δ under the take place	<u>ns</u> : anything that doesn't take part in the reaction, but is needed to ction happen. Reaction condition information is placed above the arrow. Two common ones are: he arrow means that heat is required in order for the reaction to e.
	A chemica reaction t	al formula under the arrow usually indicates the solvent that the akes place in.
	For example, the	equation:
		$2 H_2 (g) + O_2 (g) \xrightarrow{\Delta} 2 H_2 O(\ell) + heat$
	is equivalent to th	e following statement:
	"Two molec heated to p	cules of hydrogen gas and 1 molecule of oxygen gas were roduce 2 molecules of liquid water and heat."

Big Ideas	Details	Unit: Chemical Reactions
	Homework Problems	
	Write each of the following chemical equations in words.	
	1. 2 H ₂ (g) + O ₂ (g) \longrightarrow 2 H ₂ O (ℓ) + 572 kJ	
	2. $CaC_2(cr) + 2 H_2O(\ell) \longrightarrow C_2H_2(g) + Ca(OH)_2(ppt)$	
	3. 2 C ₂ H ₂ (g) + 5 O ₂ (g) $\xrightarrow{\Delta}$ 4 CO ₂ (g) + 2 H ₂ O (ℓ) + 2 60	0 kJ
	4. 3 CaCl ₂ (aq) + 2 K ₃ PO ₄ (aq) \longrightarrow Ca ₃ (PO ₄) ₂ (ppt) + 6 K	Cl (aq)

		Chernical Eq.		1 ugc. 575
Big Ideas	Details			Unit: Chemical Reactions
	Write each of tl	ne following word prob	lems as a chemical e	equation.
	5. One mol hydroflu molecule	ecule of silicon dioxide oric acid to produce a r es of liquid water.	gas reacts with four nolecule of silicon te	molecules of aqueous etrafluoride gas and two
	6. Two mo of aqueo	les of aqueous potassiu ous potassium chloride	m chlorate decomp and three moles of o	ose to produce two moles oxygen gas.
	7. Four mo oxygen ያ	les of solid antimony an gas to form one mole o	re heated in the pres f solid antimony (III)	sence of three moles of oxide.
	8. When 2 25 mole liquid wa	moles of liquid octane s of oxygen gas, 16 mol ater are formed, and 10	(C ₈ H ₁₈) are burned ir es of carbon dioxide 150 kJ of heat is pro	n the presence of e gas and 18 moles of oduced.

Unit: Chemical Reactions

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to ...)

• Recognize & identify the five major classes of chemical reactions.

Success Criteria:

• Reactions are correctly identified.

Tier 2 Vocabulary: reaction

Language Objectives:

• Explain what happens in each of the types of reaction.

Notes:

Details

There are many types of chemical reactions. Five of the most common are:

synthesis: two or more reactants combine to form a single product. For example:

$$Na + Cl_2 \rightarrow NaCl$$

<u>decomposition</u>: one reactant disintegrates (decomposes) to form two or more products:

$$H_2CO_3 \rightarrow H_2O + CO_2$$

single replacement (sometimes called single displacement): atoms of one element replace atoms of another in a compound:

$$\underline{AI} + CuCl_2 \rightarrow \underline{AI}Cl_3 + Cu$$

Most of the single replacement reactions you will encounter involve metals reacting with ionic compounds. In this type of single replacement reaction, a positive ion (usually a metal) replaces the other positive ion, or a negative ion (often a non-metal) replaces the other negative ion.

Types of Chemical Reactions

Big Ideas	Details	Unit: Chemical Reactions
	double replacement (sometimes reaction): when two positive ion to form two new compounds. Fo	called a double displacement or metathesis s (or two negative ions) switch with each other or example:
	Ca	$Cl_2 + Na_2CO_3 \rightarrow CaCO_3 + NaCl$
	Ca starts out paired with Cl, and Na trade places so that Ca is now think of it as Cl and CO ₃ trading p	Na is paired with CO ₃ . In the reaction, Ca and with CO ₃ and Na is now with Cl. (Or you could laces—the result is the same.)
	<u>combustion:</u> a special kind of reaction containing only carbon and hydro form CO ₂ and H ₂ O. For example:	on in which a hydrocarbon (a compound ogen) reacts with O2 (burns, or "combusts") to
	C ₃ H ₈ (ℓ) + 5 O ₂ (g)	\rightarrow 3 CO ₂ (g) + 4 H ₂ O (g) + heat
	All flames are chemical reactions energy—produced by the reaction reactions involving hydrocarbons	. (The flame itself is the light—photons of on.) Most flames are produced by combustion s and oxygen.
	The internal combustion engine which octane (C ₈ H ₁₈) and other h producing heat. The heat makes expanding gases push the piston	n your car is a special chemical reactor, in ydrocarbons combust in a chamber (cylinder), the gases inside the cylinder expand. The , which makes the car go.
	Lise this space for summany and/or a	dditional notes:
	ose this space for summary anu/or a	

Types of Chemical Reactions

	Types of	Chemical Neactions	Page: 376
Big Ideas	Details		Unit: Chemical Reactions
	Homework Problems For each of the following chemical reactions, indicate whether the type of r		S
			nether the type of reaction
	• synthesis	 single replacement 	 combustion
	decomposition	 double replacement 	 none of the above
	1. $H_2 + O_2 \rightarrow H_2O$	13. C ₂ H ₂ +	$O_2 \rightarrow CO_2 + H_2O$
	2. $S_8 + O_2 \rightarrow SO_2$	14. H ₂ O ₂ -	\rightarrow H ₂ O + O ₂
	3. HgO \rightarrow Hg + O ₂	15. C ₇ H ₁₆	$+ O_2 \rightarrow CO_2 + H_2O$
	4. $Zn + HCl \rightarrow ZnCl_2 +$	+ H ₂ 16. SiO ₂ +	${\rm HF} \rightarrow {\rm SiF_4} + {\rm H_2O}$
	5. Na + H ₂ O \rightarrow NaOF	H + H ₂ 17. KClO ₂	\rightarrow KCl + O ₂
	$6. C_{10}H_{16} + CI_2 \rightarrow C +$	HCI 18. KCIO ₂	\rightarrow KClO ₄ + KCl
	7. $Si_2H_2 + O_2 \rightarrow SiO_2$	+ H ₂ O 19. P ₄ O ₁₀	+ $H_2O \rightarrow H_2PO_4$
	8. Fe + $O_2 \rightarrow Fe_2O_2$	20. Sb + C	$D_2 \rightarrow Sb_4O_6$
	9. $C_7H_6O_2 + O_2 \rightarrow CC$	0 ₂ + H ₂ O 21. C ₂ H ₈ +	$O_2 \rightarrow CO_2 + H_2O$
	10. $FeS_2 + O_2 \rightarrow Fe_2O_2$	2 + SO ₂ 22. Fe ₂ O ₂	+ CO \rightarrow Fe + CO ₂
	11. $Fe_2O_2 + H_2 \rightarrow Fe +$	- H ₂ O 23. PCl ₅ +	$H_2O \rightarrow HCI + H_2PO_4$
	12. K + Br ₂ \rightarrow KBr	24. H ₂ S +	$CI_2 \rightarrow S_8 + HCI$

Types of Chemical Reactions

Page: 377

Big Ideas	Details	Unit: Chemical Reactions
	25. Fe + $H_2O \rightarrow Fe_2O_4 + H_2$	38. $K_2O + H_2O \rightarrow KOH$
	26. $N_2 + H_2 \rightarrow NH_2$	39. Al + $O_2 \rightarrow Al_2O_3$
	27. $N_2 + O_2 \rightarrow N_2O$	40. Na ₂ O + H ₂ O \rightarrow NaOH + O ₂
	28. $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$	41. C + H ₂ O \rightarrow CO + H ₂
	29. SiCl ₄ + H ₂ O \rightarrow H ₄ SiO ₄ + HCl	42. $H_3AsO_4 \rightarrow As_2O_5 + H_2O$
	30. $H_2PO_4 \rightarrow H_4P_2O_7 + H_2O$	43. $AI_2(SO_4)_2 + Ca(OH)_2 \rightarrow AI(OH)_2 + CaSO_4$
	31. $CO_2 + NH_2 \rightarrow OC(NH_2)_2 + H_2O$	44. $FeCl_2 + NH_4OH \rightarrow Fe(OH)_2 + NH_4CI$
	32. $AI(OH)_2 + H_2SO_4 \rightarrow AI_2(SO_4)_2 + H_2O$	45. $Ca_3(PO_4)_2 + SiO_2 \rightarrow P_4O_{10} + CaSiO_2$
	33. $Fe_2(SO_4)_2 + KOH \rightarrow K_2SO_4 + Fe(OH)_2$	46. $N_2O_5 + H_2O \rightarrow HNO_2$
	34. $H_2SO_4 + HI \rightarrow H_2S + I_2 + H_2O$	47. AI + HCI \rightarrow AICI ₃ + H ₂
	35. Al + FeO \rightarrow Al ₂ O ₂ + Fe	48. $H_3BO_2 \rightarrow H_4B_6O_{11} + H_2O$
	36. Na ₂ CO ₂ + HCl \rightarrow NaCl + H ₂ O + CO ₂	49. Mg + N ₂ \rightarrow Mg ₂ N ₂
	$37. P_4 + O_2 \rightarrow P_2O_5$	50. NaOH + Cl ₂ \rightarrow NaCl + NaClO + H ₂ O

Predicting the Products of Chemical Reactions

Unit: Chemical Reactions

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Accurately predict the products of single replacement, double replacement, and combustion reactions.

Success Criteria:

- Cation & anion are correct for single and double replacement reactions.
- Products that are ionic compounds have correctly balanced charges.

Tier 2 Vocabulary: product, replacement

Language Objectives:

• Explain how you can tell from the reactants what the reaction is likely to be.

Notes:

Recognizing Reaction Types from the Reactants

If you are familiar with the different types of chemical reactions, you can often tell the reaction type by looking only at the reactants. Once you know the reaction type, it is relatively straightforward to predict what the products should be.

Combustion Reactions

In a combustion reaction, the reactants are always a hydrocarbon (with some unknown number of atoms C, H, and O) and oxygen (O_2). The products are always CO_2 and H_2O .

If you were asked to write a combustion reaction for C_5H_{12} , you would write:

$$C_5H_{12} + O_2 \rightarrow CO_2 + H_2O$$

Predicting the Products of Chemical Reactions Page: 380

	Fredicting the Froducts of chem	IICAI REACTIONS Page: 380			
Big Ideas	Details	Unit: Chemical Reactions			
	Single Replacement Reactions				
	Single and double replacement reactions usually involve ionic compounds (and sometimes water, which we treat as the ionic compound H ⁺ OH ⁻).				
	In a single replacement reaction, atoms of an element react with a compound, replacing the atom of the same type. Metals replace metals; non-metals replace non-metals. For example:				
	$Na + MgI_2 \rightarrow NaI + Mg$	$KBr + Cl_2 \rightarrow KCl + Br_2$			
	(Na replaces Mg.)	(Cl replaces Br.)			
	If an element reacts with a compound, you can element simply replaces the other element of t	predict the products, because the he same type.			
	For example, if you were given the problem:				
	Ca + NaCl →	?			
	Calcium is a metal, so it will replace sodium. Th chloride (CaCl ₂), and sodium will end up by itsel	um is a metal, so it will replace sodium. This means calcium will end up with ride (CaCl ₂), and sodium will end up by itself (Na). The reaction is therefore:			
	$Ca + NaCl \rightarrow CaCl_2 + Na$				
	Remember that we have to balance the charges together. This is the most common mistake beg balance the charges in the new compounds. Yo reactants as "unbalancing" the charges, which r them when you put them back together in a ne	s <u>every time</u> we put two new ions ginning students make—forgetting to ou can think of breaking apart the means you need to "re-balance" w arrangement.			
	In the example above, the Na ion (which has a - for the charges to balance, but the Ca ion (whic	+1 charge) needed only one Cl⁻ ion h has a +2 charge) needs two.			

Predicting the Products of Chemical Reactions Page: 381

Dig Idooc	Predicting the Froducts of Chemical Reactions Page 381
Big lueas	
	Double Replacement Reactions
	In a double replacement reaction, the two ions of the same type switch places, as in: $KCl + MgO \rightarrow MgCl_2 + K_2O$
	(K and Mg are trading places; in the products, K is now with O and Mg is now with Cl.)
	Notice again that we had to balance the charges. We needed only one K^+ ion with Cl^- , but we need 2 K^+ ions with O^{2-} . Similarly, Mg^{2+} needed only one O^{2-} ion, but it needs two Cl^- ions.
	If we had the problem:
	$NH_4OH + Ca_3(PO_4)_2 \rightarrow ?$
	we would swap NH_4^+ with Ca^{2+} . When we balance the charges, NH_4^+ would go with PO_4^{3-} to form $(NH_4)_3PO_4$, and Ca^{2+} would go with OH^- to form $Ca(OH)_2$. This gives the equation:
	$NH_4 OH + Ca_3(PO_4)_2 \rightarrow (NH_4)_3 PO_4 + Ca(OH)_2$
	Acid-base reactions are a type of double replacement reaction in which H^+ and OH^- ions combine to form "HOH", which is really H_2O . For example:
	$HCl + Ca(OH)_2 \rightarrow CaCl_2 + H_2O$

Predicting the Products of Chemical Reactions Page: 382

Details	Unit: Chemical Reactions			
Homework Problems				
Predict the products for each of the following single replacement, double replacement and combustion reactions. (<i>Don't forget to balance the charges!</i>)				
1. Ca + AlCl ₃ \rightarrow				
2. $BaCl_2 + O_2 \rightarrow$				
3. KCl +Mg(OH) ₂ \rightarrow				
4. Na ₃ PO ₄ +MgSO ₄ \rightarrow				
5. Na + HCl \rightarrow				
6. Al + CoCl ₂ \rightarrow				
7. Pb(NO ₃) ₂ + NaI →				
8. $SiO_2 + Ca \rightarrow$				
9. Al + FeO \rightarrow				
10. Zn + HCl \rightarrow				
	Homework Predict the products for each of the followir replacement and combustion reactions. (Dot 1. $Ca + A C _3 \rightarrow$ 2. $BaCl_2 + O_2 \rightarrow$ 3. $KCl + Mg(OH)_2 \rightarrow$ 4. $Na_3PO_4 + MgSO_4 \rightarrow$ 5. $Na + HCl \rightarrow$ 6. $Al + CoCl_2 \rightarrow$ 7. $Pb(NO_3)_2 + NaI \rightarrow$ 8. $SiO_2 + Ca \rightarrow$ 9. $Al + FeO \rightarrow$ 10. $Zn + HCl \rightarrow$			

	Predicting the Products of Chemical	Reactions Page: 3	383
Big Ideas	Details	Unit: Chemical Reaction	ons
	11. Mg + NaNO ₃ →		
	12. NaOH + HNO ₃ →		
	13. Mg + CrCl ₃ →		
	14. Fe(OH) ₃ + H ₂ SO ₄ →		
	15. H ₂ S + AuCl ₃ →		
	16. $Cr_2(SO_4)_3$ + KOH →		
	17. PbCl₂ + Na₂SO₄ →		
	18. AgNO ₃ + CuCl ₂ →		
	19. BaCl ₂ + Al ₂ (SO ₄) ₃ →		
	20. TiO ₂ + HCl →		
	21. Zn + AgNO ₃ →		
	22. Al + Fe ₃ O ₄ \rightarrow		
	1		

Activity (Reactivity) Series

Unit: Chemical Reactions

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Use the activity series to predict whether or not a single replacement reaction will occur.

Success Criteria:

Details

- Prediction is correct about whether or not a reaction occurs.
- Cation & anion are correct if reaction does occur.
- Products have correctly balanced charges.

Tier 2 Vocabulary: product, replacement, activity

Language Objectives:

• Explain how you can tell using the activity series whether or not a reaction will occur.

Notes:

In the reaction between aluminum metal and copper (II) chloride:

$$AI(s) + CuCl_2(aq) \rightarrow AICl_3(aq) + Cu(s) + heat$$
(1)

the beaker got hot. This means the reaction gave off heat, which was lost to the surroundings (the water that the chemicals were dissolved in, the beaker, the air, your hand). Once the energy was given off, the chemicals didn't have enough energy to go the other direction. In other words, the reverse reaction does not happen:

$$Cu(s) + AlCl_3(aq) \rightarrow no reaction$$
 (2)

Is it possible to predict which direction the reaction will go?

For single replacement reactions, there is a list, called the <u>activity series</u>, (or reactivity series), which lists metals in order from most reactive to least, based on how much energy they give off when they lose electrons to form a positive ion. A metal that's higher on the list can replace anything lower on the list (because more energy is given off), but a metal that's lower on the list doesn't have enough energy to replace one that's higher up.

Big Ideas	Details	Act	ivity (R	eactiv	vity) Series	Page: 385 Unit: Chemical Reactions
		Activity (Reactivity) Series				
	Metal	Ion			Reacts With	Method of Extraction
	Rh	C3 Rh⁺	▲	•		
	K	K⁺			cold H ₂ O	
	Na	Na⁺			dilute acids	
	Li	Li ⁺			O ₂	
	Sr	Sr ²⁺			_	electrolysis
	Ca	Ca ²⁺				
	Mg	Mg ²⁺				
	Be	Be ²⁺				
	AI	Al ³⁺			at a a m	
	Mn	Mn ²⁺	increa	sing	steam dilute acide	
	Zn	Zn ²⁺	reacti	vity		
	Cr	Cr ³⁺	1		O_2	
	Fe	Fe ²⁺				
	Cd	Cd ²⁺				motal axida raduction
	Со	Co ²⁺				with carbon or CO-
	Ni	Ni ²⁺			dilute acids	
	Sn	Sn ²⁺			O ₂	smelting with coke
	Pb	Pb ²⁺				sincling with coke
	H ₂	H⁺				
	Cu	Cu ²⁺				
	Cu	Cu⁺			O ₂	
	Hg	Hg²⁺				
	Ag	Ag⁺				heat or physical
	Au	Au ³⁺			some strongly	extraction
	Pt	Pt²⁺			oxidizing acids	
	To answer th activity serie but copper o	he origir es. This can't rep	nal questic means alu Al (s) + Cu place alum Cu	on, notic Iminum uCl ₂ (aq) inum: (s) + AICl	e that aluminum i can replace coppe $\rightarrow AlCl_3(aq) + Cu(s)$ $_3(aq) \rightarrow no reaction$	s higher than copper on the er:) + heat on

Activity (Reactivity) Series

Big Ideas	Details	Unit: Chemical Reaction
		Homework Problems
	For eac	h of the following single replacement reactions:
	a)	Check the activity series to see whether the reaction happens.
	b)	If the reaction happens, predict the products. If the reaction does not happen, write "N.R." ("No Reaction").
	1.	$K(s) + H_2O \rightarrow$
	2.	Pb (s) + Zn(CH ₃ COO) ₂ (aq) \rightarrow
	3.	Al (s) + Fe ₂ O ₃ (s) $\xrightarrow{\Delta}$
	4.	AgNO ₃ (aq) + Ni (s) →
	5.	Ag (s) + H ₂ SO ₄ (aq) \rightarrow
	6.	NaBr (aq) + I ₂ (s) \rightarrow
	7.	Ca (s) + MgSO ₄ (aq) \rightarrow
	8.	Ca (s) + HCl (aq) \rightarrow
	9.	Mg (s) + HNO ₃ (aq) \rightarrow
	10.	$CuCl_2$ (aq) + Hg (ℓ) \rightarrow
	11.	Na (s) + H ₂ O (ℓ) \rightarrow

Unit: Chemical Reactions

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to...)

• Apply the law of definite proportions to balance chemical equations.

Success Criteria:

Details

• Equation is balanced such that there are the same number of atoms (moles) of each element on each side of the equation.

Tier 2 Vocabulary: balance

Language Objectives:

• Explain the law of definite proportions and conservation of mass and relate them to chemical equations.

Notes:

A chemical equation needs to describe the chemical formulas and relative number of molecules involved of each molecule that reacts, and each molecule that is produced.

Remember from Dalton's theory of the atom:

"Atoms are neither created nor destroyed in any chemical reaction."

Therefore, not only must we have the same kinds of atoms (same elements) on both sides of a chemical reaction, we need to have the *same number* of each kind of atom before and after the reaction.

Big Ideas	Details Unit: Chemical Reactions
	For example, consider the chemical equation:
	$S + O_2 \rightarrow SO_3$
	There are 2 oxygen atoms on the left, but 3 on the right. We can't change the formulas of the molecules that take part in the reaction, so we need to specify different numbers of each molecule to "balance" the equation.
	The easiest solution would be to split an O_2 molecule in half:
	$S + 1\frac{1}{2}O_2 \rightarrow SO_3$
	But we can't have $\frac{1}{2}$ of a molecule of O ₂ . Therefore, the smallest set of integers that give us the same number of each atom on both sides would be:
	$2S + 3O_2 \rightarrow 2SO_3$
	This works because there are 2 atoms of S and 6 atoms of O on each side of the equation ("before" and "after").
	We balanced this equation by inspection, but for more complicated equations, it helps to have a system.

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	Balancing Chemical Equations Page: 389
Big Ideas	Details Unit: Chemical Reactions
	To balance an equation, start with one element. Put coëfficients in front of the molecules that contain the element so that you have the same number on each side. Then do the same for every other element. For example, to balance the equation:
	$_N_2 + _H_2 \rightarrow _NH_3$
	we need to figure out the coëfficients that go in the blanks. We can start by balancing any element we want, so let's start with nitrogen (N). The smallest numbers that we can use to balance atoms of N are a "1" in front of N_2 and a 2 in front of N_3 . This gives us:
	$\underline{1} \mathrm{N}_2 + \underline{-} \mathrm{H}_2 \longrightarrow \underline{2} \mathrm{N} \mathrm{H}_3$
	Now we have 2 atoms of N on each side, so N is balanced. Next, we move on to H. We already have a "2" in front of NH_3 , which means we have 6 atoms of H on the right side. To get 6 atoms of H on the left side, we need a "3" in front of H_2 . This gives us the equation:
	$\underline{1}N_2+\underline{3}H_2\longrightarrow\underline{2}NH_3$
	We have coefficients in front of all of the products and reactants, so the equation is balanced. For the final form of the equation, we leave out any coefficient that is "1". (This is just like algebra—we would write "x" instead of "1x".) This gives us:
	$N_2 + 3 H_2 \rightarrow 2 NH_3$
	This equation was equally easy to balance regardless of whether we started with N or H, but for more complicated equations, making good decisions about what order to balance the elements in can make a huge difference.

Big Ideas	Details	Unit: Chemical Reactions			
		Strategy for Balancing Equations			
	1.	Figure out which elements to balance First, Middle, and Last. We will refer to this method as the "FML" method. [*] Always start by deciding which elements to save for last:			
		 <u>Last</u>: any element that appears by itself (anywhere in the equation) 			
		• <u>First</u> : elements that appear in only one molecule on each side (if you haven't already saved them for last).			
		<u>Middle</u> : every element that's not already last or first.			
	2.	Start with any element on the "First" list. Add coëfficients to make it balance.			
	3.	Pick another element. (Work your way through the "First," then "Middle," then "Last" lists.) Start with elements that already have at least one coëfficient, but still need at least one.			
	4.	Repeat step #3 until everything is balanced.			
	Notes	5:			
	•	Polyatomic ions usually stay together.			
	•	If you end up with a fraction, write it in temporarily, then multiply <i>all</i> of your coëfficients by the denominator of the fraction to get back to whole numbers.			
	* Now y	ou can say to yourself, "I have to balance this equation? FML."			

Big Ideas	Details		0		Unit: Chemical Reactions
	Example:				
			H ₂ S0	$D_4 + HI \longrightarrow H_2S + I_2 + H_2O$	
	1. Ma	ke lists:			
	1. 2.	<u>Last</u> : <u>First</u> :	I S, O	I appears by itself as I ₂ S only appears in H ₂ SO ₄ right; O only appears ir on the right.	(on the right) on the left and H_2S on the h H_2SO_4 on the left and H_2O
	3.	<u>Middle</u> :		H appears in two places	each on the left and right.
	2. Bal	ance "First"	elements	(S & O; the order doesn't	matter):
	a.	Let's start v the smalles	with S. Ne st pair that	ither H_2SO_4 nor H_2S has a works for both: 1 of eacl	coëfficient, so we choose h:
			<u>1</u> H ₂ SO ₄	$+ _HI \rightarrow \underline{1}_{H_2}S + _I_2 + _H$	₂ O
	b.	Next balan we have 4 H ₂ O to have	ce O. We a atoms of C e 4 atoms	already have a "1" in fron) on the left. This means v of O on the right.	t of H_2SO_4 , which means we need a "4" in front of
		$1 H_2SO_4 + HI \rightarrow 1 H_2S + I_2 + H_2O$			
	3. Bal (2 i ato ren	ance "Middl n the 1 H ₂ S a ms so far or naining 8 ato	e" elemen and 8 more of the left. oms of H.	ts (H). We have a total of e in the 4 H₂O), and our co This means we need an "8	f 10 H atoms on the right ëfficients only show 2 H 3″ in front of HI for the
			1 H ₂ SO ₄	$+ \underline{8} HI \rightarrow 1 H_2S + \underline{I}_2 + 4 H_2$	I ₂ O
	4. Bal mo	ance "Last" lecules of I ₂	elements (.:	(I). We have 8 atoms of I	on the left, which is 4
			1 H ₂ SO ₄	+ 8 HI \rightarrow 1 H ₂ S + <u>4</u> I ₂ + 4 H	l ₂ O
	5. For	the final an	swer, leav	e out any coëfficient of 1:	
			H_2SO_4	$+ 8 \text{ HI} \rightarrow \text{H}_2\text{S} + 4 \text{ I}_2 + 4 \text{ H}_2\text{G}$	C

Big Ideas	Details		Unit: Chemical Reaction
		Homework Problems	
		Set #1 (Easier)	
	Balance tl	he following chemical equations.	
	1.	$H_2 + O_2 \rightarrow H_2O$	
	2.	S_8 + $O_2 \rightarrow SO_3$	
	3.	$HgO \rightarrow Hg + O_2$	
	4.	$Zn + HCl \rightarrow ZnCl_2 + H_2$	
	5.	Na + $H_2O \rightarrow$ NaOH + H_2	
	6.	$C_{10}H_{16}$ + $CI_2 \rightarrow C$ + HCl	
	7.	$Si_2H_3 + O_2 \rightarrow SiO_2 + H_2O$	
	8.	$Fe + O_2 \rightarrow Fe_2O_3$	
	9.	$C_7H_6O_2 + O_2 \rightarrow CO_2 + H_2O$	

Use this space for summary and/or additional notes:

Page: 392 it: Chemical Reactions

	Page: 393
Unit: Chemical	Reactions

Big Ideas	Details		Unit: Chemical Reactior
	10.	$FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$	
	11.	$Fe_2O_3 + H_2 \rightarrow Fe + H_2O$	
	12.	$K + Br_2 \rightarrow KBr$	
	13.	$C_2H_2 + O_2 \rightarrow CO_2 + H_2O$	
	14.	$H_2O_2 \rightarrow H_2O + O_2$	
	15.	$C_7H_{16} + O_2 \rightarrow CO_2 + H_2O$	
	16.	$SiO_2 + HF \rightarrow SiF_4 + H_2O$	
	17.	$KCIO_3 \rightarrow KCI + O_2$	
	18.	$KCIO_3 \rightarrow KCIO_4 + KCI$	
	19.	$P_4O_{10} + H_2O \rightarrow H_3PO_4$	
	20.	$Sb + O_2 \rightarrow Sb_4O_6$	

Big Ideas	Details		Unit: Chemical R
		Homework Problems	
		Set #2 (More Challenging)	
	Balance t	he following chemical equations.	
	1.	$Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$	
	2.	$Ca_3P_2 + H_2O \rightarrow Ca(OH)_2 + PH_3$	
	3.	$Ca + AICI_3 \rightarrow CaCI_2 + AI$	
	4.	$H_3PO_3 \rightarrow H_3PO_4 + PH_3$	
	5.	$C_6H_6 + O_2 \rightarrow CO_2 + H_2O$	
	6.	$AI_4C_3 + H_2O \rightarrow CH_4 + AI(OH)_3$	
	7.	$Ag_2S + KCN \rightarrow KAg(CN)_2 + K_2S$	
	8.	$MgNH_4PO_4 \rightarrow Mg_2P_2O_7 + NH_3 + H_2O$	

Net Ionic Equations

Unit: Chemical Reactions

MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to...)

• Write chemical equations as net ionic equations.

Success Criteria:

- Soluble ionic compounds are dissociated.
- Insoluble ionic compounds remain together as solids.
- Spectator ions are identified and omitted from the final net ionic equation.

Tier 2 Vocabulary: net, spectator

Language Objectives:

• Review dissociation. Explain how to dissociate a compound and write its ions separately in an equation.

Notes:

<u>net ionic equation</u>: a chemical equation that shows only ions or pure substances that are changed by the reaction.

<u>spectator ion</u>: an ion that remains in solution and does not participate in a chemical reaction.

If you mixed aqueous solutions of calcium chloride $(CaCl_2(aq))$ and sodium nitrate $(NaNO_3(aq))$, you might be tempted to predict that the following (unbalanced) chemical reaction would occur:

 $CaCl_2(aq) + NaNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + NaCl(aq)$

However, recall that aqueous ions dissociate when they dissolve in water:

Symbol	What Actually Happens in H_2O	
CaCl ₂ (aq)	Ca ²⁺ (aq) + Cl⁻ (aq)	
NaNO₃ (aq)	$Na^+(aq) + NO_3^-(aq)$	
Ca(NO ₃) ₂ (aq)	$Ca^{2+}(aq) + NO_{3}^{-}(aq)$	
NaCl (aq)	Na ⁺ (aq) + Cl ⁻ (aq)	

This means that what we really have in the beaker is:

 $Ca^{2+}(aq) + Cl^{-}(aq) + Na^{+}(aq) + NO_{3}^{-}(aq) \rightarrow Ca^{2+}(aq) + NO_{3}^{-}(aq) + Na^{+}(aq) + Cl^{-}(aq)$ The above is called a detailed ionic equation.

..:

	Net Ionic Equations	Page: 396		
Big Ideas	Details	Unit: Chemical Reactions		
	In the detailed ionic equation:			
	$Ca^{2+}(aq) + Cl^{-}(aq) + Na^{+}(aq) + NO_{3}^{-}(aq) \rightarrow Ca^{2+}(aq) + NO_{3}^{-}(aq) + Na^{+}(aq) + Cl^{-}(aq)$			
	Notice that the right side and the left side contain exactly the same ions. In other words, <i>nothing has changed</i> . If no substances are changed—no chemical bonds are formed or broken—then <u>no chemical reaction has occurred</u> !			
	Now consider the reaction of aqueous silver nitrate with aqueous sodium carbonate:			
	$AgNO_3(aq) + Na_2CO_3(aq) \rightarrow Ag_2CO_3(s) +$	- NaNO ₃ (aq)		
	Notice that one of the products, silver carbonate, forms	a solid (precipitate).		
	The detailed ionic equation would look like this:			
	$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3$	(s) + $Na^+(aq) + NO_3^-(aq)$		
	The spectator ions (ions that remain unchanged by the I If we cross those out:	reaction) are Na ⁺ and NO ₃ ⁻ .		
	$Ag^+(aq) + N\Theta_3^-(aq) + Na^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3$	(s) + Na ⁺ (aq) + NO 3 ⁻ (aq)		
	we are left with the <u>unbalanced net ionic equation</u> for this reaction:			
	$Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3$	(s)		
	Of course, we still need to balance the equation! The <u>ba</u> would therefore be:	<u>alanced net ionic equation</u>		
	$2 \operatorname{Ag}^{+}(\operatorname{aq}) + \operatorname{CO}_{3}^{2^{-}}(\operatorname{aq}) \rightarrow \operatorname{Ag}_{2}\operatorname{CO}$	3 (S)		
	Notice that the net ionic reaction is much simpler than t because the net ionic equation leaves out everything th equation, allowing you to focus only on the details that	the full chemical equation, at does not matter in the are important.		
	If you take AP [®] Chemistry, you will be expected to write net ionic form.	e all chemical equations in		

Use this space for summary and/or additional notes:
Net Ionic Equations

Big Ideas	Details Unit	: Chemical Reactions
	Homework Problems	
	For each of the following potential double replacement reaction	IS:
	a. Predict the products. (Remember to balance the charges	5!)
	b. Use your solubility rules to write the phase after each pro If the product is soluble, write (aq) after it. If an ionic con soluble, then it precipitates; write (ppt) after it. If a produ CO_2), then write (g) after it. If a product is a pure liquid (s write (ℓ) after it.	oduct. npound is not uct is a gas (such as such as H ₂ O), then
	c. Rewrite the equation with the aqueous compounds disso	ciated (split up).
	d. Cancel (cross out) any ions that are the same (unchanged	l) on both sides.
	 e. Write and balance the net ionic equation as your final an that you have crossed out everything, write "N.R." ("No F 	swer. If it turns out Reaction") instead.
	1. $Na_2CO_3 (aq) + CaCl_2 (aq) \rightarrow$	
	2. (NH₄)₃PO₄ (aq) + NaOH (aq) →	
	3. Ba(C ₂ H ₃ O ₂) ₂ (aq) + K ₃ PO ₄ (aq) \rightarrow	
	4. Ca(MnO₄)₂ (aq) + KOH (aq) →	

Net Ionic Equations

Big Ideas	Details		Unit: Chemical Reactions
	5.	$AICI_3 (aq) + H_3PO_4 (aq) \rightarrow$	
	6.	CaSO₄ (aq) + KMnO₄ (aq) →	
	7.	NaN₃ (aq) + Ca(NO₃)₂ (aq) →	
	8.	$Sr(NO_3)_2$ (aq) + K ₂ Cr ₂ O ₇ (aq) \rightarrow	
	9.	NaClO₃ (aq) + MgSO₄ (aq) →	
	10.	Na ₃ BO ₃ (aq) + ZnSO ₄ (aq) →	

Big Ideas	Details Unit: Oxidation & Reduction
	Introduction: Oxidation & Reduction
	Unit: Oxidation & Reduction
	Topics covered in this chapter:
	Oxidation-Reduction (REDOX) Reactions
	Electrochemical Cells407
	Standards addressed in this chapter:
	Massachusetts Curriculum Frameworks & Science Practices (2016):
	HS-PS1-10(MA) Use an oxidation-reduction reaction model to predict products of reactions given the reactants, and to communicate the reaction models using a representation that shows electron transfer (redox). Use oxidation numbers to account for how electrons are redistributed in redox processes used in devices that generate electricity or systems that prevent corrosion.

Unit: Oxidation & Reduction

MA Curriculum Frameworks (2016): HS-PS1-10(MA)

Mastery Objective(s): (Students will be able to...)

- Assign oxidation numbers.
- Write and balance equations for simple REDOX reactions.

Success Criteria:

Details

- Oxidation numbers agree with published/accepted values.
- Oxidation numbers add up to zero for compounds and to the charge for ions.
- Balanced REDOX reactions have the same number of each type of atom and the same number of electrons transferred on each side.

Tier 2 Vocabulary: reduce

Language Objectives:

• Explain electron transfer. Explain the charge that an ion gets when electrons are added or removed.

Notes:

oxidation-reduction reaction (REDOX reaction): a reaction in which one or more electrons are transferred from one atom to another.

In the 1700s, oxidation meant that an atom was combined with oxygen, and was therefore "oxidized". For example:

$$2 \text{Cu} + \text{O}_2 \rightarrow 2 \text{CuO}$$

In this reaction, oxygen is more electronegative than copper, so oxygen took electrons away from copper. This means that the copper (which was "oxidized" by oxygen) lost two electrons and ended up with a +2 charge. Oxygen gained two electrons and ended up with a -2 charge. As a result, the term "oxidation" has come to mean "losing electrons".

Also, in the 1700s, scientists found that if they heated the CuO (in which copper has a +2 charge), they ended up with copper metal (with a charge of zero), and the weight was reduced. Thus copper was said to be "reduced". As a result, the term "reduction" has come to mean "gaining electrons".

Big Ideas	Details	Unit: Oxidation & Reduction
	oxidation: the loss of one or more electrons by an atom	in a chemical reaction
	reduction: the gain of one or more electrons by an atom	n in a chemical reaction.
	Stupid Mnemonics: There are two popular mnemonics and reduction, one "Democratic" and one "Rep	for remembering oxidation ublican".
	LEO the lion says 'GER' ("Democratic" mnemonic invol LEO stands for "Loss of Electrons is Oxidation" and (Electrons is Reduction"	ving endangered species): GER stands for "Gain of
	OIL RIG ("Republican" mnemonic involving oil companie Involves Loss (of electrons)", and RIG stands for "Re electrons)."	es): OIL stands for "Oxidation duction Involves Gain (of
	In a redox reaction, at least one element is oxidized, and is reduced. An element cannot be oxidized in a chemica element is reduced, and vice-versa. (After all, the electr somewhere, and they have to go somewhere.)	d at least one other element <i>I reaction unless some other</i> ons have to come from
	All chemical reactions in which an element becomes parversa, are redox reactions. This includes all single replacombustion reactions, and many synthesis and decomp chemists generally classify a reaction as a redox reaction the energy of the reaction comes from electron transfer	rt of a compound, or vice- cement reactions, osition reactions. However, n only when most or all of r.

Big Ideas	Details	Unit: Oxidation & Reduction
	Oxidation Number	S
	An oxidation number is a measure of how "oxidized" ar neither oxidized nor reduced, so it has an oxidation nur	atom is. An element is nber of zero.
	 An element that has lost electrons (oxidized) gets number, equal to the number of electrons it has l 	a positive oxidation ost.
	 An element that has gained electrons (reduced) g number, equal to the number of electrons it has g 	ets a negative oxidation gained.
	Therefore:	
	when an element is oxidized, the oxidation numb	er increases.
	 When an element is reduced, the oxidation numb (decreases). 	er is also reduced
	oxidation number (or "oxidation state"): the charge tha compound if all bonds were completely ionic and ev or ion had a charge.	t an atom would have in a very atom in the compound
	L Use this space for summary and/or additional notes:	

Big Ideas	Details	Unit: Oxidation & Reduction
		Assigning Oxidation Numbers
	 The c The c The c polya In a c Th ne to Al Al Al Al Al Ca 	exidation number of a pure element is 0. (Even if it's diatomic.) exidation numbers in a compound add up to 0 exidation number of an ion is its charge. (Oxidation numbers in a tomic ion add up to the charge of the polyatomic ion.) ompound or polyatomic ion: we most electronegative element (the last one in the formula) has a egative oxidation number equal to the number of electrons it would need fill its valent shell. I other atoms have positive oxidation numbers. uorine is always -1 . eygen is always -2 except in the peroxide ion ($O_2^{2^-}$) and in OF ₂ . eydrogen is always +1 except in metal hydrides (such as NaH). kali (group 1) metals are always +1. kaline Earth (group 2) metals are always +2. is always +3, Zn is always +2, and Ag is always +1. hlculate other elements from the above.
	Comple D	roblem.
	What are t	roblem: he oxidation numbers of each element in the compound Na ₂ HPO $_4$:
	• Na ₂ H	PO ₄ is an ionic compound made of the ions Na ⁺ and HPO ₄ ²⁻
	• The N	Ja^+ ion has a charge of +1, so the oxidation number of Na is +1.
	• The F and C	4^{2-} ion has a charge of -2. This means the oxidation numbers of H, P, D must add up to -2.
	o 0	- −2. There are 4 O atoms. (4)(−2) = −8
	0 H	is +1.
	○ If ox	the oxidation numbers for the O atoms add up to −8 and H is +1, then the idation number for P must be +5 so the total can add up to −2.

Big Ideas	Details	Unit: Oxidation & Reduction
		Balancing REDOX Reactions
	To fully bal	ance a redox reaction, you must balance:
	AtcEleTot	oms (as you would in a regular equation) ctrons lost/gained al charge
	Often, redo balancing t electrons a reduction h	ex reactions are shown and balanced as net ionic equations. In this case, hem is often a simple matter of making sure that the same number of re produced by the oxidation half-reaction and consumed by the half-reaction.
	For exampl	e, consider the unbalanced net ionic equation:
		$AI^{0}(s) + Zn^{2+}(aq) \rightarrow AI^{3+}(aq) + Zn^{0}(s)$
	In this reac The <u>atoms</u> produces 3	tion, Al is oxidized from Al ⁰ to Al ⁺³ , and Zn is reduced from Zn ⁺² to Zn ⁰ . appear balanced, but Zn ²⁺ needs only 2 electrons to form Zn ⁰ , but Al ⁰ electrons when oxidized to Al ³⁺ .
	The two ha	If-reactions are: Oxidation: $AI^0 \rightarrow AI^{3+} + 3e^-$ Reduction: $Zn^{2+} + 2e^- \rightarrow Zn^0$
	To balance second one	the electrons, we need to multiply the first half-reaction by 2, and the by 3, giving:
		$2 \operatorname{Al}^{0} \longrightarrow 2 \operatorname{Al}^{3+} + 6 \operatorname{e}^{-}$ $3 \operatorname{Zn}^{2+} + 6 \operatorname{e}^{-} \longrightarrow 3 \operatorname{Zn}^{0}$
	If we comb on both sid	ine these and cancel the electrons (because we have the same number es), we get the balanced net ionic equation:
		$2 \text{Al}^{0}(s) + 3 \text{Zn}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Zn}^{0}(s)$
	REDOX read reactions a participate scope of th	ctions can get a lot more complicated, especially when acid-base re also taking place and the water that the ions are dissolved in s in the reaction. Balancing complex REDOX reactions is beyond the is course, but is covered in AP [®] Chemistry.

Big Ideas	Details	Unit: Oxidation & Reduction
		Homework Problems
	For each of element.	the following compounds or ions, write the oxidation number of each
	1.	FeO
	2.	N ₂
	3.	KMnO₄
	4.	NH₄OH
	5	
	5.	
	6.	Cr ₂ O ₇ ²⁻

Use this space for summary and/or additional notes:

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	Oxidation-Reduction (REDOX) Reactions Page: 406
Big Ideas	Details Unit: Oxidation & Reduction
	Balance the following (redox) equations. A superscript "0" indicates a pure element.
	7. $7n^0 + Tl^+ \rightarrow Tl^0 + 7n^{2+}$
	7. 20 10 20
	8. $Li^0 + Cr^{3+} \rightarrow Cr^0 + Li^+$
	9. $K^0 + Mg^{2+} \rightarrow K^+ + Mg^0$

Unit: Oxidation & Reduction

MA Curriculum Frameworks (2016): HS-PS1-10(MA)

Mastery Objective(s): (Students will be able to...)

• Explain how an electrochemical cell (such as a battery) works.

Success Criteria:

- Explanations account for each of the parts of the electrochemical cell.
- Tier 2 Vocabulary: battery, bridge

Language Objectives:

• Explain how a battery works.

Notes:

Details

<u>electrochemistry</u>: using chemical (redox) reactions to produce electricity or *vice-versa*. In an electrochemical reaction, oxidation and reduction reactions occur in separate containers, and the electrons that travel from one reaction to the other pass through an electric circuit.

<u>galvanic cell</u>: (also called a voltaic cell) a chemical apparatus that uses an electrochemical reaction to produce electricity. (A battery is a type of galvanic cell.)



Electrochemical Cells

Big Ideas	Details	Unit: Oxidation & Reduction
	<u>electrolytic cell</u> : a cell similar to a galvanic cel spontaneous, and electricity is used to ad reaction occur. (Electrolysis of water is a	l, except that the reaction is not d the energy needed to make the n example.)
	<u>electrode</u> : a solid metal strip where either ox strips also conduct the electrons into or o	idation or reduction occurs. The metal of the electric circuit.
	anode: the negatively (–) charged electrode.	At the anode:
	 Oxidation happens. (Atoms from the ar These metal ions become part of the so The electrons produced by oxidation metathode: the positively (+) charged electrode 	node are oxidized to positive ions.) Iution. (<i>I.e.,</i> the anode loses mass.) ove up the wire into the electric circuit. . At the cathode:
	 Reduction happens. (Ions from the solution atoms.) 	ution are reduced to neutral metal
	 These metal ions become part of the ca The electrons needed for reduction mo wire and into the cathode. 	thode. (<i>I.e.,</i> the cathode gains mass.) ve from the electric circuit through the
	Note that in physics, electric "current" is defin particle would move. This means that the "cu from the electrons.	ed to be the direction that a <u>positive</u> prrent" flows in the <u>opposite</u> direction
	salt bridge: a salt solution that is connected t provides ions for the two half-cells to kee are not allowed to balance, opposite char the reaction would stop.) The salt solutio take part in the reactions at the cathode o	o both half cells. The salt bridge p the charges balanced. (If the charges ges would build up in both cells and on must be made of ions that do not or anode. (KNO ₃ is commonly used.)
	<u>electroplating</u> : using an electrolytic cell to ad cathode is attached to the object to be el- reduces metal ions from the solution, whi	d a layer of metal to something. The ectroplated. An electric current ich are deposited onto the object.

Introduction: Stoichiometry

Unit: Stoichiometry

Topics covered in this chapter:

Stoichiometry	410
Stoichiometry: Mass-Mass Problems	415
Limiting Reactant	419
Percent Yield	427
Marathon Problems	430

Standards addressed in this chapter:

Massachusetts Curriculum Frameworks & Science Practices (2016):

HS-PS1-7 Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.

Unit: Stoichiometry

Details

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to...)

• Solve mole-mole stoichiometry problems.

Success Criteria:

- For each compound in the chemical equation, the ratio of the coëfficients is the same as the ratio of the moles.
- Solutions have the correct quantities substituted for the correct variables.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: mole, coëfficient

Language Objectives:

• Explain how the coefficients in a chemical equation are like the numbers in a pre-algebra "input-output machine."

Notes:

stoichiometry: measurement of how much of each reactant is used and how much of each product is produced in a chemical reaction.

stoichiometry problem: a chemistry problem in which you are given a balanced chemical equation and the quantity of one compound, and you are asked to find the quantity of another compound produced or consumed in the same equation.

For example, in the chemical reaction:

 $3 \text{ CaCl}_2 + 2 \text{ Na}_3\text{PO}_4 \rightarrow 6 \text{ NaCl} + \text{Ca}_3(\text{PO}_4)_2$

3 molecules of $CaCl_2$ would produce 1 molecule of $Ca_3(PO_4)_2$. Because a mole is always the same number of molecules, this means 3 moles of $CaCl_2$ produces 1 mole of $Ca_3(PO_4)_2$.

Stoichiometry is simply the process of using the coëfficients in a balanced chemical equation to convert from moles of one compound to moles of another.



Use this space for summary and/or additional notes:



Big Ideas	Details	Unit: Stoichiometry
		Homework Problems
	1.	Determine how much of each product would be made when 4.0 mol of Pb(NO ₃) ₂ decomposes in the reaction: 2 Pb(NO ₃) ₂ \rightarrow 2 PbO + 4 NO ₂ + O ₂
	2.	Answer: 4.0 mol PbO; 8.0 mol NO ₂ ; 2.0 mol O ₂ Determine how much of each product would be made when 1.33 mol of
		Ca ₃ P ₂ reacts with excess water in the reaction: Ca ₃ P ₂ + 6 H ₂ O \rightarrow 3 Ca(OH) ₂ + 2 PH ₃
		Answer: $3.99 \text{ mol Ca}(OH)_2$; 2.66 mol PH_3
	3.	Determine how much AlCl ₃ would you need to completely react with 1.5 mol Ca in the reaction: 3 Ca + 2 AlCl ₃ \rightarrow 3 CaCl ₂ + 2 Al
		Answer: 1.0 mol AlCl₃
	4.	Determine how much of each product would be made when 1.50 mol H_3PO_3 decomposes in the reaction: $4 H_3PO_3 \rightarrow 3 H_3PO_4 + PH_3$
		Answer: 1.13 mol H ₃ PO ₄ ; 0.375 mol PH ₃

		Scolemonieury	1 ugc. +1+
Big Ideas	Details		Unit: Stoichiometry
	5.	Determine how many moles of KCl would be produced to K and excess Cl_2 in the reaction: 2 K + $Cl_2 \rightarrow$ 2 KCl	uced from 0.175 mol of
		Answer: 0.175 mol KCl	
	6.	Determine how many moles of Na ₂ O would be red 0.275 mol of NaOH in the reaction: Na ₂ O + H ₂ O \rightarrow	luired to produce ▶ 2 NaOH
		Answer: 0.138 mol Na ₂ O	
	7.	Determine how many moles of O_2 will be produced NaClO ₃ in the reaction: 2 NaClO ₃ \rightarrow 2 NaCl + 3 O_2	d by 8.75 mol of
		Answer: 26.3 mol O ₂	
	8.	Determine how many moles of NaCl are produced reaction when 45.4 L of O_2 are produced at S.T.P. i 2 NaCl $O_3 \rightarrow$ 2 NaCl + 3 O_2 (<i>Hint: you will need to convert</i> 45.4 L <i>of gas at S.T</i>	in the following n the reaction: . <i>P. into moles first.</i>)
		Answer: 1.33 mol NaCl	

Unit: Stoichiometry

Details

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to...)

• Solve stoichiometry problems that require mole conversions.

Success Criteria:

- Conversions between moles and other quantities are set up and executed correctly.
- For each compound in the chemical equation, the ratio of the coëfficients is the same as the ratio of the moles.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: mole, coëfficient

Language Objectives:

• Explain the order of operations: convert to moles, do stoichiometry, convert from moles to desired units.

Notes:

stoichiometry: measurement of how much of each reactant is used and how much of each product is produced in a chemical reaction.

Remember that stoichiometry has to be done in moles.

- If you are given amounts in any other unit, you need to convert to moles before doing stoichiometry.
- If your answer needs to be in another unit, you need to convert after doing stoichiometry.

<u>mass-mass problem</u>: a stoichiometry problem that requires mole conversions from mass of a reactant to moles, and then moles of a product back to mass.

Note that there are many other similar problems that would work the same way *e.g.,* from volume of a gas (using the ideal gas law) to moles, from volume of a liquid that has a certain concentration (in $\frac{\text{mol}}{1}$) to moles, *etc.*

Big Ideas	Details	Unit: Stoichiometry
	Sample Problem:	
	How many grams of copper metal would be produced from 13.5 excess copper chloride solution in the chemical reaction:	g of aluminum and
	2 Al (s) + 3 CuCl ₂ (aq) \rightarrow 2 AlCl ₃ (aq) + 3 Cu (s)	
	Strategy:	
	1. Convert grams of Al to moles.	
	2. Use stoichiometry to convert moles of Al to moles of Cu.	
	3. Convert moles of Cu to grams.	
	Setup:	
	$\frac{13.5 \text{gAt}}{1000 \text{ cm}} \times \frac{1 \text{motAt}}{1000 \text{ cm}} \times \frac{3 \text{motCu}}{1000 \text{ cm}} \times \frac{63.5 \text{ gCu}}{1000 \text{ cm}}$	
	1 27.0gAt 2motAl 1motCu	
	Answer:	
	$\frac{(13.5)(3)(63.5)}{(27.0)(2)} = \frac{2572}{54.0} = 47.6 \mathrm{g}\mathrm{Cu}$	
	theoretical yield: the amount of a product you could make base calculations, assuming that at least one of the reactants is co	d on stoichiometry ompletely used up.
	excess: having more of a reactant than is needed. This means si "enough that you don't have to worry about using it all up." problems in which this is not the case in the next section ("Li starting on page 419.)	mply that there is We will see miting Reactant,"
	By this point in this course, you have undoubtedly figured out th challenging problems you will encounter are created by stringing sequence of easy problems until it becomes hard to keep track o doing. Stoichiometry is easy (once you get the hang of it). Mole easy (assuming you've got the hang of them). Combining the tw it's just a sequence of easy problems.	at most of the g together a f what you're conversions are o looks hard, but

Stoichiometry: Mass-Mass Problems

Big Ideas	Details	Unit: Stoichiometry
		Homework Problems
	1.	In the chemical reaction:
		$2 \text{ K} + \text{Cl}_2 \rightarrow 2 \text{ KCl}$
		 a. How many <i>moles</i> of KCl (F.W. 74.55^g/_{mol}) would be produced from 2.50 g of K and excess Cl₂?
		Answer: 0.0639 mol KCl
		b. How many grams of KCI would be produced?
		Answer: A 76 g KCl
	2.	In the chemical reaction:
		$Na_2O + H_2O \rightarrow 2 NaOH$
		a. If 124 g of Na ₂ O (F.W. 61.98 $\frac{g}{mol}$) is reacted with excess H ₂ O, how
		many grams of NaOH (F.W. $40.00 \frac{g}{mol}$) will be made?
		Answer: 160. g NaOH
		b If instead you wanted to make 100 g of NaOH how many grams of
		Na ₂ O would you need?
		Answer: 77.5 g NaOH

Stoichiometry: Mass-Mass Problems

Big Ideas	Details	Unit: Stoichiometry
	3.	In the decomposition reaction:
		$2 \text{ NaClO}_3 \rightarrow 2 \text{ NaCl} + 3 \text{ O}_2$
		If you reacted 26.6 g of NaClO ₃ (F.W. 106.44 $\frac{g}{mol}$), what volume of O ₂ would
		you make at a pressure of 1.03 atm and a temperature of 30 °C? (<i>Hint: Use PV = nRT.</i>)
		Answer: 9.06 L O ₂ (g)
	4.	Given the precipitation reaction:
		$3 \text{ CaCl}_2 (aq) + 2 \text{ Na}_3 \text{PO}_4 (s) \rightarrow \text{Ca}_3 (\text{PO}_4)_2 (\text{ppt}) + 6 \text{ NaCl} (aq)$
		If you added an excess of powdered Na ₃ PO ₄ to 100. mL of an $0.200 \frac{\text{mol}}{1}$
		solution of CaCl ₂ , how many grams of precipitate would form? (Assume that all of the Ca ₃ (PO ₄) ₂ precipitates, and that all of the Na ₃ PO ₄ dissolves.)
		Answer: 2.07 g Ca ₃ (PO ₄) ₂ (ppt)
	5.	How many grams of precipitate would form if 94.6 g of $FeCl_3 \cdot 6 H_2O$ crystals were added to an aqueous solution containing an excess of Na_2SiO_3 ? (<i>Hint: you will need to predict the products and balance the equation in order to do the stoichiometry.</i>)
		Answer: 50.5 g Fe ₂ (SiO ₃) ₃ (ppt)

Details		Unit: Stoichiometr
	Limiting F	Reactant
Unit: Stoichiometr	Ъ	
MA Curriculum F	rameworks (2016): HS	S-PS1-7
Mastery Objectiv	re(s): (Students will be	able to)
 Identify the I 	imiting reactant in a stole	chiometry problem.
 Perform stoic reactant. 	chiometry calculations in	a problem that involves a limiting
 Determine the 	ne amount(s) of the non-	limiting reactant(s) left over.
Success Criteria:		
 Limiting read 	tant correctly identified	
 Stoichiometr compound ir 	y calculations performed the desired units).	l correctly (correct amount of desired
 Algebra and Fier 2 Vocabulary 	rounding to appropriate /: limiting	number of significant figures is correct.
Language Object	ives:	
• Explain why	a chemical reaction runs	out of something.
Notes:		
Q: What happens v	vhen a chemical reaction	runs out of something?
A: The reaction sto	ps.	
A reaction in which) you run out of somethir	ng is called a <u>l</u> imiting reactant problem.
The reactant that y because running o	ou run out of is called th ut of it is what limits how	e <u>limiting reactant</u> (or limiting reagent) / much product you can make.

Use this space for summary and/or additional notes:

Big Ideas

Consider the following reaction:

Details

Big Ideas



Suppose you have 51 cans and four six-pack rings. There are two possibilities:

- 1. We use up all of the cans. (Situation A)
- 2. We use up all of the six-pack rings. (Situation B)

	Six-Packs	Six-Pack Rings	Cans	Situation
X	could make 8.5	need at least 8.5	Have 51	А
\odot	could make 4	Have 4	need at least 24	В

As you can see, we have only enough of <u>both</u> reactants to make 4 six-packs. Once we have made 4 six-packs, we have used up all of the six-pack rings, and we cannot make any more.

This means six-pack rings are the limiting reactant, and we use all of them.

We used up 24 cans (the non-limiting reactant), and we had 27 cans left over.

On the other hand, suppose we have only 15 cans and 4 six-pack rings. Again, there are two possibilities:

- 1. We use up all of the cans. (Situation C)
- 2. We use up all of the six-pack rings. (Situation D)

	Six-Packs	Six-Pack Rings	Cans	Situation
0	could make 2.5	need at least 2.5	have 15	С
X	could make 6	have 4	need at least 24	D

This time, we can make 2.5 six-packs and then we run out of cans, so now the limiting reactant is cans. If we make 2.5 six-packs, we would use 2.5 six-pack rings, which means we would have 1.5 six-pack rings left over.

	Linning Reactant	Page: 421
Big Ideas	Details	Unit: Stoichiometry
	The secret to solving limiting reactant problems is to do a stoi each reactant (using the factor-label method) to see how muc products you make if you used it all up. The limiting reactant make the least amount of product (gets used up first).	chiometry problem on h of one of the is the one that can
	Steps for Solving Limiting Reactant	Problems
	 Does the problem give you amounts for more than on it's not a limiting reactant problem.) 	e reactant?(If not,
	2. Convert the amount of each reactant to moles.	
	3. For each of the reactants, use stoichiometry to figure of the products you could make if you used the reacta	out how much of one int up.
	 The reactant that can make the <u>least</u> amount of produused up first—it is the <i>limiting reactant</i>. 	uct is the one that gets
	5. Perform <i>all</i> of your stoichiometry calculations using the the limiting reactant that you identified in step 4 above	ne number of moles of ve.
	 If the problem asks how much of one of the non-limit over, use the moles of the limiting reactant to find ou the other reactant got used up. Subtract this number started with to find out how much is left over. 	ing reactants is left t how many moles of from the moles you
	If the problem is asking for a quantity other than mole convert from moles to the desired unit.	es (such as grams),

Big Ideas	Details	Unit: Stoichiometry
	Sample Problem:	
	Q: Given the following reaction:	
	$16 \text{ Cu} + \text{S}_8 \rightarrow 8 \text{ Cu}_2 \text{S}$	
	If we had 27.5 moles of copper and 1.4 moles of S_8 , how make?	nuch Cu₂S would we
	A: Determine how many moles of Cu ₂ S we could make from	each reactant:
	$\frac{27.5 \text{ mol Cu}}{1} \times \frac{8 \text{ mol Cu}_2 \text{S}}{16 \text{ mol Cu}} = 13.75 \text{ mol Cu}_2 \text{S}$	- more
	$\frac{1.4 \operatorname{mol} S_8}{1} \times \frac{8 \operatorname{mol} Cu_2 S}{1 \operatorname{mol} S_8} = 11.2 \operatorname{mol} Cu_2 S$	- less
	We can make 11.2 mol of Cu_2S and then we run out of S_8 limiting reactant.	$_{3}$. This means S ₈ is the
	Q: How much of the non-limiting reactant would be left over	?
	A: S_8 was limiting, and we had 1.4 moles of it. We need to fing ot used up.	nd out how much Cu
	$\frac{1.4 \operatorname{mol} S_8}{1} \times \frac{16 \operatorname{mol} Cu}{1 \operatorname{mol} S_8} = 22.4 \operatorname{mol} Cu$	₂ S
	Now we subtract to find how much was left:	
	27.5 mol – 22.4 mol = 5.1 mol Cu left o	ver

Big Ideas	Details	5	Unit: Stoichiometry
		Homework Problems	
		Set #1: Scaffolded	
	1. Coi	nsider the reaction: 2 Si ₂ H ₂ + 5 O ₂ \rightarrow 4 SiO ₂ + 2 H ₂ O	
	a	. If you had 8 mol Si_2H_2 , how many moles of O_2 wo above reaction?	uld you need for the
	h	Answer: 20 moles O_2 If you had 15 mol O_2 how many moles of SiaHa w	ould you need for
		the above reaction?	
		Answer: 6 mol Si ₂ H ₂	
	C	limiting?	ant would be
		Answer: O ₂	
	с	. How many moles of the non-limiting reactant wo	uld be left over?
		Answer: 2 mol Si ₂ H ₂ left over	
	e	 What is the theoretical yield of SiO₂, in moles? (<i>I.e</i> of SiO₂ would you make?) 	e., how many moles
		Answer: 12 mol SiO ₂	

Big Ideas	Details		Unit: Stoichiometry
	2.	Cons	ider the reaction: 3 Ca (s) + 2 AlCl ₃ (aq) \rightarrow 3 CaCl ₂ (aq) + 2 Al (s)
		a.	If you had 6 mol Ca, how many moles of AlCl₃ would you need for the above reaction?
			Answer: 4 mol AICI ₃
		b.	If you had 8 mol AlCl ₃ , how many moles of Ca would you need for the above reaction?
			Answer: 12 mol Ca
		C.	If you had 6 mol Ca and 8 mol AlCl3, which reactant would be limiting?
		d.	Answer: Ca How many moles of the non-limiting reactant would be left over?
			Answer: 4 mol AlCl₃ left over
		e.	What is the theoretical yield of $CaCl_2$, in moles? (<i>I.e.</i> , how many moles of $CaCl_2$ would you make?)
			Answer: 6 mol CaCl ₂

Big Ideas	Details	Unit: Stoichiometry
		Homework Problems
		Set #2: Unscaffolded
	1.	How many moles of H ₂ O would be produced if 3.5 mol H ₂ react with 1.5 mol O ₂ in the reaction: $2 H_2 + O_2 \rightarrow 2 H_2O$
		(Note: because amounts were given for both reactants, this is a limiting reactant problem.)
	2	Answer: 3.0 mol H_2O
	Ζ.	$S_8 + O_2 \rightarrow SO_3$
		Which reactant is limiting, and how much of the other reactant would be left over?
	3.	Answer: O ₂ is limiting; there will be3.7 mol S ₈ left over. 325 g of H ₂ O is poured onto a 450. g block of sodium metal. The equation for this reaction is: $2 \text{ Na} + 2 \text{ H}_2\text{O} \rightarrow 2 \text{ Na}\text{OH} + \text{H}_2$ a. What is the limiting reactant?
		 Answer: H₂O b. If the reaction temperature is 227 °C (500. K) at a pressure of 1 atm, how many liters of H₂ gas are produced? (<i>Hint: find the moles of H₂ produced and use the ideal gas law to calculate the volume.</i>)
		Answer: 369 L H ₂

Big Ideas	Details		-	Unit: Stoichiometry
	4.	5.00 g 2	Zn are reacted with 100. mL of 1.00 M H0	Cl in the reaction:
			Zn (s) + 2 HCl (aq) \rightarrow ZnCl ₂ (a	q) + H ₂ (g)
		a.	Determine which reactant is limiting.	
			Answer: HCl	
		b.	Determine the number of <i>arams</i> of 7nC	la that will be produced.
			Answer: 6.82 g ZnCl ₂	
		C.	If the reaction conditions are 177 °C (real Kelvin) and 1 atm pressure determine	<i>member to convert to</i>
			gas that will be produced.	
			Answer: 1851 Ha	
		Ь	Determine the mass in grams of the nor	n-limiting reactant that will
		u.	be left over.	
			Answer: $1.72 \times 7n$	
			Answer: 1.75 g Zn	

Percent Yield

Unit: Stoichiometry

Details

MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to...)

• Calculate the percent yield of a reaction.

Success Criteria:

- Theoretical yield calculated correctly using stoichiometry calculations.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: yield

Language Objectives:

• Explain how to turn fractions into percentages.

Notes:

theoretical yield: the amount of a product predicted, based only on stoichiometry calculations.

actual yield: the actual amount of product recovered in the laboratory.

<u>percent yield</u>: the amount of product recovered, expressed as a percentage of the theoretical yield.

When you do a stoichiometry calculation, the answer to the question "how much product should be produced" is the theoretical yield.

Actual yield depends on several factors. Many reactions do not go to completion, but instead reach an equilibrium condition where the amount of reactants and products is constant. Sometimes it is not possible to recover all of the product because of challenges associated with separating it from the other reactants and products. *Etc.*

Because of these factors, the actual yield is determined by performing the reaction in a laboratory and measuring the amount of product you got.

Once you have the actual and theoretical yield numbers, the percent yield is:

 $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \!=\! \text{percent yield}$

Percent Yield

Big Ideas	Details Unit: Stoichiometry
	Sample Problem:
	Q: Suppose you perform the reaction:
	3 Al (s) + 3 CuCl ₂ (aq) \rightarrow 2 AlCl ₃ (aq) + 3 Cu (s)
	If you start with 9.0 g of AI and you recover 28 g of Cu, what was your percent yield?
	A: First, calculate the theoretical (predicted) yield of Cu using stoichiometry:
	$\frac{9.0 \text{ g Al}}{1} \times \frac{1 \text{ mol Al}}{27.0 \text{ g Al}} \times \frac{3 \text{ mol Cu}}{2 \text{ mol Al}} \times \frac{63.6 \text{ g Cu}}{1 \text{ mol Cu}} = 31.8 \text{ g Cu}$
	Then calculate: $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \text{percent yield}$:
	$\frac{28 \text{ g Curecovered}}{31.8 \text{ g Cupredicted}} = 0.88 \times 100 = 88 \%$
	Note that the percent yield cannot exceed 100 %. Conservation of mass tells us that if we had enough aluminum and copper chloride in the above reaction to make 31.8 g of copper metal, there's no way we could actually make more than that. If you calculate a percent yield greater than 100 %, you should:
	 Double-check your calculations to make sure you didn't make a mistake.
	 Look for other compounds that might have gotten into the product that you measured. For example:
	 If you collect a precipitate on a piece of filter paper, it will be with with everything else that was in the beaker. Even if you evaporate the water, it will leave behind other compounds that had been dissolved. Also, if you let the product dry in the air on a humid day, the compound could be hygroscopic and/or form a hydrate.
	 If you collect a gas, the most common way is to use a eudiometer (gas collection tube) that starts filled with water, and to have the gas displace the water. However, there will also be water vapor in the tube, so you need to account for the water based on the vapor pressure of water at the temperature the gas was collected.

Percent Yield

Big Ideas	Details	Unit: Stoichiometry		
		Homework Problems		
	In order to isolate percent yield problems, these questions refer to the mass- stoichiometry homework problems starting on page 417.			
	1.	In problem #1, part b on page 417, suppose that 3.85 g KCl was recovered in the lab. What was the percent yield?		
		Answer:		
	2.	In problem #2, part a on page 417, suppose that 125 g of NaOH was recovered. What was the percent yield?		
		Answer:		
	3.	In problem #3 on page 418 suppose that 10 L of O ₂ was recovered. a. What was the percent yield?		
		Answer:		
		b. What might have happened in the lab that could account for the fact that you got a percent yield higher than 100 %?		
	4.	In problem #4 on page 418 suppose you started with 3.00 g of precipitate, which was still wet from the solution in the reaction. You let it dry in the lab until the mass stopped changing, and you recorded it to be 2.50 g. Is this a good answer? If not, what else could you do?		
	Use thi	s space for summary and/or additional notes:		

Marathon Problems

Unit: Stoichiometry

Details

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-7, HS-PS2-8(MA) Mastery Objective(s): (Students will be able to...)

• Solve challenging problems that combine several aspects of chemistry.

Success Criteria:

- Problems correctly utilize strategies from various topics throughout the year.
- Solutions use the equation appropriate for the information given.
- Solutions have the correct quantities substituted for the correct variables.
- Algebra and rounding to appropriate number of significant figures is correct.

Language Objectives:

• Explain what each part of each problem is asking and which topic it relates to.

Notes:

These are intentionally challenging problems that relate topics we studied throughout the year, including gas laws, solutions & concentration, solubility, naming compounds and writing formulas, predicting products, activity series, balancing equations, stoichiometry, limiting reactant, and percent yield.

A couple of the answers are provided so you can check your work at key points in the process, but you're on your own for the rest!

Marathon Problems

Big Ideas	Details		Unit: Stoichiometry
			Homework Problems
	1.	0.75 L o chlorid	of 2.5 M sodium phosphate is mixed with 1.25 L of 2.0 M calcium e.
		a.	Predict the products, write and balance the chemical equation.
		b.	Use your solubility rules to determine whether a chemical reaction happens based on whether a precipitate forms. (<i>If no reaction happens, you may skip the rest of this question. Hint: a reaction <u>does</u> happen. ^(C))</i>
		c.	Calculate the number of moles of each reactant and determine which one is limiting.
		d.	How many grams of the precipitate are produced?
			Answer: 260 g
		e.	If 150.g of precipitate was recovered on the filter paper, what is the percent yield?
		f.	If one of the products remains in solution (<i>Hint: it does.</i>), what is its concentration in $\frac{\text{mol}}{L}$?
			(Hint: you will need to add the volumes of the two solutions that you started with to find the total volume.)

Marathon Problems

Big Ideas	Details		Unit: Stoichiometry
	2.	In a lal oxygen This co same a	poratory experiment 115 g of sulfur (S ₈) was reacted with 89.6 L of gas (O ₂) at S.T.P., in a synthesis reaction, producing only compound X. mpound contains 50% sulfur by mass, and its empirical formula is the s its molecular formula.
		a.	Use percent composition data to determine the chemical formula of compound <i>X</i> .
		b.	Write a balanced chemical equation for the reaction.
		C.	Which of the reactants was limiting?
		d.	What is the theoretical yield of compound <i>X</i> , in grams?
			Answer: 229 g
		e.	If 189 g of compound <i>X</i> was actually recovered, what was the percent yield of <i>X</i> ?
			Answer: 82.4 %

Use this space for summary and/or additional notes:
Big Ideas	Introduction: Thermochemistry (Heat) Page: 433 Details Unit: Thermochemistry (Heat)
	Introduction: Thermochemistry (Heat)
	Unit: Thermochemistry (Heat)
	Topics covered in this chapter:
	Heat & Temperature434
	Specific Heat Capacity & Calorimetry437
	Phase Changes & Heating Curves
	Thermodynamics452
	Enthalpy of Formation459
	Heat of Reaction
	Bond Energies468
	Standards addressed in this chapter:
	Massachusetts Curriculum Frameworks & Science Practices (2016):
	 HS-PS1-3 Cite evidence to relate physical properties of substances at the bulk scale to spatial arrangements, movement, and strength of electrostatic forces among ions, small molecules, or regions of large molecules in the substances. Make arguments to account for how compositional and structural differences in molecules result in different types of intermolecular or intramolecular interactions.
	HS-PS1-4 Develop a model to illustrate the energy transferred during an exothermic or endothermic chemical reaction based on the bond energy difference between bonds broken (absorption of energy) and bonds formed (release of energy).
	HS-PS3-4b Provide evidence from informational text or available data to illustrate that the transfer of energy during a chemical reaction in a closed system involves changes in energy dispersal (enthalpy entropy* change) and heat content (entropy enthalpy change) while assuming the overall energy in the system is conserved.
	 * The MA 2016 Curriculum Frameworks reversed the parenthetical references to entropy and enthalpy. I have corrected them in these notes.
	Use this space for summary and/or additional notes:

Big Ideas	Details Unit: Thermochemistry (Heat)
	Heat & Temperature
	Unit: Thermochemistry (Heat)
	MA Curriculum Frameworks (2016): HS-PS1-4
	Mastery Objective(s): (Students will be able to)
	• Explain the difference between heat and temperature.
	 Describe what is happening at the molecular level when a system is in thermal equilibrium.
	Success Criteria:
	 Explanation accounts for total energy as well as direction of energy flow ("driving force").
	 Description accounts for and relates macroscopic observations to microscopic phenomena.
	Tier 2 Vocabulary: heat, temperature
	Language Objectives:
	 Explain the difference between heat and temperature.
	Notes:
	<u>heat</u> : energy that can be transferred when moving atoms or molecules collide with each other.
	temperature: a measure of the average kinetic energy of the particles (atoms or molecules) of a system.
	thermometer: a device that measures temperature, most often via thermal expansion/contraction of a liquid or solid.

Heat & Temperature



Heat & Temperature

			rage. 430
Big Ideas	Details		Unit: Thermochemistry (Heat)
	As an analogy, heat tra	insfer is a lot like flowing wate	r.
	 Water flows from a higher temperature 	om a higher elevation to a low erature to a lower one.	er one, just like heat flows from
	 The total energy height of the water molecula total heat (energy) 	gy of the water going over the vaterfall (the average gravitati es) and the total mass of wate ergy) contained in an object de temperature.	waterfall depends on both the onal potential energy of the r going over it. Similarly, the pends on both the mass of the

Specific Heat Capacity & Calorimetry

Unit: Thermochemistry (Heat)

MA Curriculum Frameworks (2016): HS-PS2-6, HS-PS3-1

Mastery Objective(s): (Students will be able to...)

- Calculate the heat transferred when an object with a known specific heat capacity is heated.
- Perform calculations related to calorimetry.
- Describe what is happening at the molecular level when a system is in thermal equilibrium.

Success Criteria:

- Variables are correctly identified and substituted correctly into the correct equations.
- Algebra is correct and rounding to appropriate number of significant figures is reasonable.

Tier 2 Vocabulary: heat, specific heat capacity, coffee cup calorimeter

Language Objectives:

- Explain what the specific heat capacity of a substance measures.
- Explain how heat is transferred between one substance and another.

Labs, Activities & Demonstrations:

• Calorimetry lab.

Notes:

Different objects have different abilities to hold heat. For example, if you enjoy pizza, you may have noticed that the sauce holds much more heat (and burns your mouth much more readily) than the cheese or the crust.

The amount of heat that a given mass of a substance can hold is based on its specific heat capacity.

Specific Heat Capacity & Calori

	Specific Heat Capacity & Calorimetry Page: 438
Big Ideas	Details Unit: Thermochemistry (Heat)
	specific heat capacity (C): a measure of the amount of heat required per gram of a substance to produce a specific temperature change in the substance.
	C_p : specific heat capacity, measured at constant pressure. For gases, this means the measurement was taken allowing the gas to expand as it was heated.
	C _v : specific heat capacity, measured at constant volume. For gases, this means the measurement was made in a sealed container, allowing the pressure to rise as the gas was heated.
	For solids and liquids, $C_p \approx C_v$ because the pressure and volume change very little as they are heated. For gases, $C_p > C_v$ (always). For ideal gases, $C_p - C_v = R$, where R is a constant known as "the gas constant."
	When there is a choice, C_p is more commonly used than C_v because it is easier to measure. When dealing with solids and liquids, most physicists just use C for specific heat capacity and don't worry about the distinction.
	Calculating Heat from a Temperature Change
	The amount of heat gained or lost when an object changes temperature is given by the equation:
	$Q = mC\Delta T$
	where:
	Q = heat (J or kJ)
	m = mass (g or kg)
	$C = \text{specific heat capacity } \left(\frac{J}{g^{\circ}C}\right)$
	ΔT = temperature change (K or °C)
	Note that $1_{g \cdot C}^{J} \equiv 1_{kg \cdot C}^{kJ} \equiv 1_{g \cdot C}^{J}$.
	You need to be careful with the units. If the mass is given in kilograms (kg), your specific heat capacity will have units of $\frac{kJ}{kg^{\circ}C}$ and the heat energy will come out in
	kilojoules (kJ). If mass is given in grams, you will use units of $\frac{J}{g^{.\circ}C}$ and the heat energy will come out in joules (J).
	Use this space for summary and/or additional notes:

Specific Heat Capacity & Calorimetry

Big Ideas

Unit: Ther	mochemistry	(Heat)
------------	-------------	--------

Specific Heat Capacities of Some Substances			
Substance	Specific Heat Capacity $\left(\frac{J}{g^{.\circ}C}\right)$	Substance	Specific Heat Capacity $\left(\frac{J}{g^{.\circ}C}\right)$
water at 20 °C	4.181	aluminum	0.897
ethylene glycol	2.460	glass	0.84
(anti-freeze)	2.400	iron	0.450
ice at −10 °C	2.080	copper	0.385
steam at 100 °C	2.11	brass	0.380
steam at 130 °C	1.99	silver	0.233
vegetable oil	2.00	lead	0.160
air	1.012	gold	0.129

Calorimetry

calorimetry: the measurement of heat flow

In a calorimetry experiment, heat flow is calculated by measuring the mass and temperature change of an object and applying the specific heat capacity equation.

<u>calorimeter</u>: an insulated container for performing calorimetry experiments.

<u>coffee cup calorimeter</u>: a calorimeter that is only an insulated container—it does not include a thermal mass (such as a mass of water). It is usually made of styrofoam, and is often nothing more than a styrofoam coffee cup.

<u>bomb calorimeter</u>: a calorimeter for measuring the heat produced by a chemical reaction. A bomb calorimeter is a double-wall metal container with water between the layers of metal. The heat from the chemical reaction makes the temperature of the water increase. Because the mass and specific heat of the calorimeter (water and metal) are known, the heat produced by the reaction can be calculated from the increase in temperature of the water.

It has a great name, but a bomb calorimeter doesn't involve actually blowing anything up. $\textcircled{\mbox{$\odot$}}$

Specific Heat Capacity & Calorimetry

	Specific field capacity & calofificity Fage. 440
Big Ideas	Details Unit: Thermochemistry (Heat)
	Solving Coffee Cup Calorimetry Problems
	Most coffee cup calorimetry problems involve placing a hot object in contact with a colder one. Many of them involve placing a hot piece of metal into cold water.
	To solve the problems, assume that both objects end up at the same temperature.
	If we decide that heat gained (going into a substance) by each object that is getting hotter is positive, and heat lost (coming out of a substance) by every substance that is getting colder is negative, then the basic equation is:
	Heat Lost + Heat Gained = Change in Thermal Energy $\sum Q_{n+1} \sum Q_{n+2} = AQ$
	$\sum Q_{lost} + \sum Q_{gained} = \Delta Q$
	If the calorimeter is insulated, then no heat is gained or lost by the entire system (which means $\Delta Q = 0$).
	If we have two substances (#1 and #2), one of which is getting hotter and the other of which is getting colder, then our equation becomes:
	Heat Lost + Heat Gained = Change in Thermal Energy
	$\sum Q_{i,j} + \sum Q_{i,j} = \Delta Q = 0$
	$\sum_{\alpha_{lost}} \sqrt{\frac{1}{2}} \frac{\alpha_{gained}}{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{\alpha_{$
	$m_1c_1\Delta t_1 + m_2c_2\Delta t_2 = 0$
	In this example, ΔT_1 would be negative and ΔT_2 would be positive.
	To solve a calorimetry problem, there are six quantities that you need: the two masses, the two specific heat capacities, and the two temperature changes. (You might be given initial and final temperatures for either or both, in which case you'll need to subtract. Remember that if the temperature increases, ΔT is positive, and if the temperature decreases, ΔT is negative.) The problem will usually give you all but one of these and you will need to find the missing one.
	If you need to find the final temperature, use $\Delta T = T_f - T_i$ on each side. You will
	have both T_i numbers, so the only variable left will be T_f . (The algebra is straightforward, but ugly.)

Specific Heat Capacity & Calorimetry Page: 441 ..

Big Ideas	Details	Unit: Thermochemistry (Heat)
	Sample Problems:	
	Q: An 0.050 kg block of aluminum is heated and place 0.100 kg of water at 20. °C. If the final temperature what temperature was the aluminum heated?	ced in a calorimeter containing are of the water was 30.°C, to
	A: To solve the problem, we need to look up the special uminum and water in the table on page 439. The aluminum is $0.898 \frac{J}{g^{-C}}$, and the specific heat capa. We also need to realize that we are looking for the aluminum. ΔT is always <i>final</i> – <i>initial</i> , which mean the aluminum starts out at a higher temperature, number, which is what we want.)	ecific heat capacities for the specific heat capacity of acity for water is $4.181 \frac{J}{g^{\circ}C}$. The initial temperature of the tans $\Delta T_{AI} = 30 - T_{i,AI}$. (Because , this will give us a negative
	$m_{Al}C_{Al}\Delta T_{Al} + m_{w}C_{w}\Delta T_{w}$ (0.050)(0.897)(30 - T _i) + (0.100)(4.181)(30 - 2 0.0449(30 - T _i) + 4.181 1.3455 - 0.0449T _i + 4.18 5.520	= 0 = 0 = 0 81 = 0 65 = 0.0449 T _i T _i = $\frac{5.5265}{0.0449}$ = 123.2 °C
	 Q: An 0.025 kg block of copper at 95°C is dropped in 0.075 kg of water at 25°C. What is the final temp 	to a calorimeter containing erature?
	A: We solve this problem the same way. The specifi $0.385 \frac{J}{g^{\circ}C}$, and $\Delta T_{Cu} = T_f - 95$ and $\Delta T_w = T_f - 25$. two places. The algebra will be even uglier, but it Algebra 1 problem:	ic heat capacity for copper is This means <i>T_f</i> will appear in t's still a straightforward
	$m_{Cu}C_{Cu}\Delta T_{Cu} + m_{w}C_{w}\Delta T_{v}$ $(0.025)(0.385)(T_{f} - 95_{i}) + (0.075)(4.181)(T_{i})$ $0.009625(T_{f} - 95) + 0.3138(T_{f} - 0.009625)(95) + 0.3136T_{f} - (0.3138)(T_{f} - 0.009625)(95) + 0.3138T_{f} - (0.3138)(T_{f} - 0.009625)(T_{f} - 0.9144 + 0.3138)(T_{f} - 0.009625)(T_{f} - 0.9144 + 0.009625)(T_{f} - 0.9144)(T_{f} - 0.009625)(T_{f} - 0.009625)(T_{f} - 0.9144)(T_{f} - 0.009625)(T_{f} - 0.9144)(T_{f} - 0.009625)(T_{f} - 0.009625)(T_{f} - 0.9144)(T_{f} - 0.009625)(T_{f} - 0.009625)(T_{f} - 0.9144)(T_{f} - 0.009625)(T_{f} - 0.009625)(T_{$	= 0 $T_{f} - 25) = 0$ 25) = 0 138)(25) = 0 -7.845 = 0 $3234 T_{f} = 8.759$ $T_{f} = \frac{8.759}{0.3234} = 27 \text{°C}$
		0.3234

Specific Heat Capacity & Calorimetry Page: 442

	Specific Heat Capacity & Calorimetry Page: 44
Big Ideas	Details Unit: Thermochemistry (Hea
	Homework Problems
	You will need to look up specific heat capacities in Table Z. Selected Properties of the Elements, starting on page 516.
	 375 kJ of heat is added to a 25.0 kg granite rock. If the temperature increases by 19.0 °C, what is the specific heat capacity of granite?
	Answer: 0.790 Jg.°C
	2. A 0.040 kg block of copper at 95 °C is placed in 0.105 kg of water at an unknown temperature. After equilibrium is reached, the final temperature is 24 °C. What was the initial temperature of the water?
	Answer: 21.5 °C
	3. A sample of metal with a specific heat capacity of 0.50 ^J / _{g.°C} is heated to 98 ° and then placed in an 0.055 kg sample of water at 22 °C. When equilibrium is reached, the final temperature is 35 °C. What was the mass of the metal?
	Answer: 0.0948 kg

Big Ideas	Details	Unit: Thermochemistry (Heat)
	4.	A 0.280 kg sample of a metal with a specific heat capacity of $0.430 \frac{1}{\pi^{3}}$ is
		heated to 97.5 °C then placed in an 0.0452 kg sample of water at 31.2 °C.
		What is the final temperature of the metal and the water?
		Answer: 57 °C

Use this space for summary and/or additional notes:

Specific Heat Capacity & Calorimetry Page: 443

Chemistry 1

Unit: Thermochemistry (Heat)

MA Curriculum Frameworks (2016): HS-PS1-2, HS-PS1-3, HS-PS3-4b

Mastery Objective(s): (Students will be able to...)

• Determine the amount of heat required for all of the phase changes that occur over a given temperature range.

Success Criteria:

- Variables are correctly identified and substituted correctly into the correct equations.
- Algebra is correct and rounding to appropriate number of significant figures is reasonable.

Tier 2 Vocabulary: specific heat capacity, heating curve

Language Objectives:

• Explain what the heat is used for in each step of a heating curve.

Labs, Activities & Demonstrations:

- Evaporation from washcloth.
- Fire & ice (latent heat of paraffin).

Notes:

<u>phase</u>: a term that relates to how rigidly the atoms or molecules in a substance are connected.

solid: molecules are rigidly connected. A solid has a definite shape and volume.

<u>liquid</u>: molecules are loosely connected—bonds are continuously forming and breaking. A liquid has a definite volume, but not a definite shape.

- gas: molecules are not connected. A gas has neither a definite shape nor a definite volume. Gases will expand to fill whatever space they occupy.
- <u>plasma</u>: the system has enough heat to remove electrons from atoms, which means the system is comprised of particles with rapidly changing charges.
- <u>phase change</u>: when an object or substance changes from one phase to another through gaining or losing heat.

0.1 ~ .. _

	Phase Changes & Heating Curves Page: 445
Big Ideas	Details Unit: Thermochemistry (Heat)
	Breaking bonds requires energy. Forming bonds releases energy. This is true for the bonds that hold a solid or liquid together as well as for chemical bonds (regardless of what previous teachers may have told you!)
	<i>I.e.,</i> you need to add energy to turn a solid to a liquid (melt it), or to turn a liquid to a gas (boil it). Energy is released when a gas condenses or a liquid freezes. (<i>E.g.,</i> ice in your ice tray needs to give off heat in order to freeze. Your freezer needs to remove that heat in order to make this happen.)
	The reason evaporation causes cooling is because the system (the water) needs to absorb heat from its surroundings (<i>e.g.</i> , your body) in order to make the change from a liquid to a gas (vapor). When the water absorbs heat from you and evaporates, you have less heat, which means you have cooled off.
	Calculating the Heat of Phase Changes
	<u>heat of fusion</u> (Δ <i>H_{fus}</i>) (sometimes called "latent heat" or "latent heat of fusion"): the amount of heat required to melt one kilogram of a substance. This is also the heat released when one kilogram of a liquid substance freezes. For example, the heat of fusion of water is $334 \frac{1}{g}$. The heat required to melt a sample of
	water is therefore:
	$Q = m\Delta H_{fus} = m(334 \frac{J}{g})$
	<u>heat of vaporization</u> (ΔH_{vap}): the amount of heat required to vaporize (boil) one kilogram of a substance. This is also the heat released when one kilogram of a
	gas condenses. For example, the heat of vaporization of water is $2260 \frac{J}{g}$. The
	heat required to boil a sample of water is therefore:
	$Q = m\Delta H_{vap} = m (2260 \frac{J}{g})$



In the "solid" portion of the curve, the sample is solid water (ice). As heat is added, the temperature increases. The specific heat capacity of ice is $2.11 \frac{J}{g^{.}C}$, so the heat required is:

$$Q_{solid} = mC\Delta T = m(2.11\frac{J}{g^{\circ}C})\Delta T$$

In the "melting" portion of the curve, the sample is a mixture of ice and water. As heat is added, the ice melts, but the temperature remains at 0°C until all of the ice is melted. The heat of fusion of ice is $334 \frac{1}{g}$, so the heat required is:

$$Q_{melt} = m\Delta H_{fus} = m(334 \frac{J}{g})$$

In the "liquid" portion of the curve, the sample is liquid water. As heat is added, the temperature increases. The specific heat capacity of liquid water is $4.181 \frac{J}{g^{.}C}$, so the heat required is:

$$Q_{liquid} = mC\Delta T = m(4.181\frac{J}{g^{\circ}C})\Delta T$$

	
Big Ideas	Details Unit: Thermochemistry (Heat)
	In the "boiling" portion of the curve, the sample is a mixture of water and water
	vapor (steam). As neat is added, the water bolls, but the temperature remains at
	100 C until all of the water has bolled. The heat of vaporization of water is $2260\frac{1}{g}$,
	so the heat required is:
	$Q_{melt} = m\Delta H_{vap} = m(2260 \frac{J}{g})$
	In the "gas" portion of the curve, the sample is water vapor (steam). As heat is added, the temperature increases. The specific heat capacity of steam is
	approximately $2.08 \frac{J}{g^{\circ}C}$. (This is at 100 °C; the specific heat capacity of steam
	decreases as the temperature increases.) The heat required is:
	$Q_{gas} = mC\Delta T = m(2.08 \frac{J}{g^{\circ}C})\Delta T$
	Steps for Solving Heating Curve Problems
	A heating curve problem is a problem in which a substance is heated across a temperature range that passes through the melting and/or boiling point of the substance, which means the problem includes heating or cooling steps and melting/freezing or boiling/condensing steps.
	 Sketch the heating curve for the substance over the temperature range in question. Be sure to include the melting and boiling steps as well as the heating steps.
	 From your sketch, determine whether the temperature range in the problem passes through the melting and/or boiling point of the substance.
	3. Split the problem into:
	a. Heating (or cooling) steps within each temperature range.
	b. Melting or boiling (or freezing or condensing) steps.
	4. Find the heat required for each step.
	a. For the heating/cooling steps, use the equation $Q = mC\Delta T$.
	b. For melting/freezing steps, use the equation $Q - m\Delta H$.
	b. For methal steps, use the equation $Q = m \Delta r_{fus}$.
	c. For boiling/condensing steps, use the equation $Q = m \Delta H_{vap}$.
	5. Add the values of <i>Q</i> from each step to find the total.
	Use this space for summary and/or additional notes:

Big Ideas	Details	Unit: Thermochemistry (Heat)
	Sample Problem	
	Q: How much heat would it take to raise the temperator -25.0 °C to +130.0 °C?	erature of 15.0 g of H_2O from
	A: The H_2O starts out as ice. We need to:	
	1. Heat the ice from −25.0 °C to its melting po	int (0 °C).
	2. Melt the ice.	0°C to 100°C)
	4. Boil the water.	0 C (0 100 C).
	5. Heat the steam from 100 °C to 130 °C.	
	6. Add up the heat for each step to find the to	otal.
	heat solid: $Q_1 = mC\Delta T = (15)(2.11)(25) = 791.25$.	I
	melt the ice: $Q_2 = m \Delta H_{fus} = (15)(334) = 5010 \text{ J}$	
	heat liquid: $Q_3 = mC\Delta T = (15)(4.181)(100) = 627$	0 1
	boil: $Q_4 = m \Delta H_{vap} = (15)(2260) = 33900\text{J}$	
	heat gas: $Q_5 = mC\Delta T = (15)(2.08)(30) = 936 \text{ J}$	
	$Q = Q_1 + Q_2 + Q_3 + Q_4 + Q_5$ $Q = 791 + 5010 + 6270 + 33900$	+936=46910 J

/es Page: 449 Unit: Thermochemistry (Heat)

Big Ideas

Homework Problems

For the following problems, use data from the following table:

	C (sol.) $\left(\frac{kJ}{kg^{\circ}C}\right)$	M.P. (°C)	ΔH_{fus} $\left(rac{kJ}{kg} ight)$	C (liq) (<u>kJ</u> kg.°C)	B.P. (°C)	ΔH_{vap} $\left(rac{kJ}{kg} ight)$	C _p (gas) (<u>kJ</u> (kg.°C)
water	2.11	0	334	4.18	100	2260	2.08*
potassium	0.560	62	61.4	1.070	760	2025	0.671
mercury	0.142	-39	11.3	0.140	357	293	0.104
silver	0.217	962	111	0.318	2212	2360	—

*Note that because of the volume change from heating, the specific heat capacity of gases, C_p , increases with increasing temperature.

- 1. A 0.0250 kg sample of water is heated from -40.0° C to 150. °C.
 - a. Sketch the heating curve for the above process. Label the starting temperature, melting point, boiling point, and final temperature on the *y*-axis.

b. Calculate the heat required for each step of the heating curve, and the total heat required.

Answer: 80.01 kJ

Phase Changes	& Heating	Curves
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Big Ideas	Details		Unit: Thermochemistry (Heat)
	2.	A 0.085	5 kg sample of mercury is heated from 25°C to 500. °C.
		a.	Sketch the heating curve for the above process. Label the starting temperature, melting point, boiling point, and final temperature on the <i>y</i> -axis.
		b.	Calculate the heat required for each step of the heating curve, and the total heat required.
			Anguar: 20.12 kl

Phase Changes & Heating Curves Details Unit: Thermochemistry (Heat) 3. A 0.045 kg block of silver at a temperature of 22°C is heated with 20.0 kJ of energy.

a. Calculate the total heat required by calculating the heat for each step until the entire 20.0 kJ is accounted for.

liquid, gas) of the silver at that temperature?

b. What is the final temperature and what is the physical state (solid,

Answer: liquid, 1369°C

Use this space for summary and/or additional notes:

Big Ideas

Big Ideas

Thermodynamics

Unit: Thermochemistry (Heat)

MA Curriculum Frameworks (2016): HS-PS3-4b

Mastery Objective(s): (Students will be able to...)

• Explain the laws of thermodynamics.

• Apply the laws of thermodynamics conceptually to hypothetical situations.

Success Criteria:

- Explanations account for enthalpy and entropy differences.
- Explanations account for the conservation of energy.

Tier 2 Vocabulary: system, free

Language Objectives:

• Explain the laws of thermodynamics.

Notes:

In the universe, energy is what "makes things happen". In chemistry, we use the following terms to distinguish between the molecules we are talking about vs. the ones we aren't:

system: the collection of molecules under consideration for a given situation.

surroundings: everything that is not part of the system

E.g., for a bunch of chemicals in a beaker, the chemicals would be the <u>system</u>, and the beaker, the air in the room, and everything else would be the <u>surroundings</u>.

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Details

Heat Flow

We generally use the variable *Q* to represent heat.

Heat flow is always represented in relation to the system.

Heat Flow	Sign of Q	System	Surroundings
from surroundings	+ (positive)	gains heat	lose heat
to system		(gets warmer)	(get colder)
from system	– (negative)	loses heat	gain heat
to surroundings		(gets colder)	(get hotter)

A positive value of *Q* means heat is flowing *into* the system. Because the heat is transferred from the molecules outside the system to the molecules in the system, the temperature of the system increases, and the temperature of the surroundings decreases.

A negative value of *Q* means heat is flowing <u>out of</u> the system. Because the heat is transferred from the molecules in the system to the molecules outside the system, the temperature of the system decreases, and the temperature of the surroundings increases.

<u>endothermic reaction</u>: a chemical reaction in which heat energy in the system is used to make the reaction proceed. This causes the system to get colder, which then causes heat to flow into the system from the surroundings. (Positive value of Q.)

<u>exothermic reaction</u>: a chemical reaction in which heat energy is released as the reaction proceeds. This causes the system to get hotter, which then causes heat to flow out of the system into the surroundings. (Negative value of Q.)

This is confusing for most people.

If you add sodium hydrogen carbonate (baking soda) to a strong acid, the solution will get colder, and the beaker will feel cold to the touch. The reaction took heat away from the solution. What makes this an endothermic reaction is that the system (the solution) is now colder than the surroundings, which means heat flows from the surroundings into the system to warm it back up.

If you add sodium hydroxide (lye) to a strong acid, the solution will get hotter, and the beaker will feel hot to the touch. The reaction released heat into the solution. What makes this an exothermic reaction is that the system (the solution) is now hotter than the surroundings, which means heat flows from the system into the surroundings to heat up the surroundings (and therefore cool off the system).

Big Ideas	Details	Unit: Thermochemistry (Heat)
	A quick and simple way to think of this is b water is the system and your hand is part of the glass, your hand gets colder because he surroundings) into the system (the ice wate heat, and the surroundings are losing heat example.	y thinking of a glass of ice water. The ice of the surroundings. When you pick up eat is flowing from your hand (the er). This means the system is gaining . The value of <i>Q</i> would be positive in this
	thermal equilibrium: when all of the partic kinetic energy (temperature).	les in a system have the same average
	When a system is at thermal equilibriu collisions between particles still transfe temperature of the particles in the syst can measure with a thermometer—is r	m, no net heat is transferred. (<i>I.e.,</i> er energy back and forth, but the average tem—the macroscopic quantity that we not changing.)

Big Ideas	Details Unit: Thermochemistry (Heat)
	Usable vs. Non-Usable Energy
	In any system, some of the energy
	enthalpy: the total usable thermal (heat) energy of a system.
	<u>entropy</u> : thermal energy that exists in a system, but cannot be transferred to other molecules or objects.
	Energy can only be transferred in a reasonable amount of time if there is enough of an energy difference between the particles that are supplying the energy and the particles receiving it. For example, if your body temperature is 37 °C (98.6 °F) and you jump into water that is 10 °C (50 °F), you will lose heat very quickly. However, if you climb into a 36.9 °C spa ("hot tub"), you can stay in for an hour and your body temperature will be unaffected.
	Entropy is therefore the energy that has been dispersed ("lost") into the surroundings, because there is so little difference between its heat content and the heat content of the surroundings that the energy cannot be transferred to other particles.
	However, in many cases it is possible to reduce the entropy of a system by adding energy.
	For example, if you open a bottle of ethyl acetate (the solvent in nail polish), the molecules will gradually diffuse into the room, and the entire room will smell like nail polish. The entropy of the gases in the room has increased, because the molecules of ethyl acetate (and the energy they contain) are more spread than they were before.
	You could build a machine to pump all of the air in the room through a -40 °C condenser. This would condense 99 % of the ethyl acetate to a liquid, which you could then pour back into the bottle. This would indeed reduce entropy, but you would have to use much more energy to condense the ethyl acetate than the amount of energy you would recover in the form of reduced entropy.
	Use this space for summary and/or additional notes:

Big Ideas	Details Unit: Thermochemistry (Heat)
	Thermodynamics
	In chemistry, we categorize energy in the following ways:
	<u>kinetic energy</u> (<i>K</i>): the energy contained in the particles due to motion of the particles or sub-particles (atoms, molecules, atomic nuclei, electrons, <i>etc.</i>) The equation for kinetic energy is $K = \frac{1}{2}mv^2$, so kinetic energy depends on the masses and velocities (speeds) of the particles.
	<u>temperature</u> (<i>T</i>): the average kinetic energy of the particles in a system. You can think of it as "heat per molecule". Increasing the temperature makes the molecules move faster; decreasing the temperature makes the molecules move more slowly.
	internal energy (U): the total kinetic energy [*] of the molecules in a system.
	Internal energy depends on the average kinetic energy of the molecules (temperature, <i>T</i>) and the number of molecules (or number of moles of molecules, <i>n</i>). For the units to work out so the energy comes out in units of joules, we multiply by the gas constant, giving the equation:
	$U = \frac{3}{2}nRT$
	2
	work (W): the energy that a gas can transfer by its pressure causing a change in volume. Work can be done on the surroundings (the gas decreases its energy), or work can be done by the surroundings on the gas (the gas increases its energy).
	$W = P\Delta V$
	enthalpy (H): the total usable energy of a system, which includes its internal energy plus energy in the form of work that could be transferred between the system and surroundings.
	H = U + PV
	* Actually, internal energy also includes microscopic potential energy, which the energy of
	(internal induced electric or magnetic dipole moment, stress-strain energy due to deformation of solids, <i>etc.</i>). Microscopic potential energy is beyond the scope of this course.

Use this space for summary and/or additional notes:

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Big Ideas	Details	Uni	t: Thermochemistry (Heat)		
	<u>entropy</u> (S): energy that e molecules or objects.	exists in a system, but cannot be transferred to other			
	free energy: the energy of a system that is "free" (available) to do work. (Sometimes called the usable energy of the system.)				
	<u>Helmholtz free energy</u> (A): the usable energy of a system with a constant volume. Named after the German physicist Hermann von Helmholtz. (Useful in physics; rarely used in chemistry.)				
		A = U - TS			
	<u>Gibbs free energy</u> (G): the Named after the Amer (Useful in chemistry; r	e usable energy of a system with rican physicist J. Willard Gibbs. arely used in physics.)	constant pressure.		
		G = H - TS			
	In chemistry, Gibbs free er what extent) a chemical re	nergy is the most useful in predic eaction will happen.	ting whether (and to		
	A chemical reaction is <u>spontaneous</u> if the change in Gibbs free energy is negative $(\Delta G < 0)$, <i>i.e.</i> , if the reaction <i>releases</i> energy and the total energy of the products is less than the total energy of the reactants.				
	A chemical reaction is <u>non-spontaneous</u> if the change in Gibbs free energy is positive $(\Delta G > 0)$, <i>i.e.</i> , if the reaction <i>consumes</i> energy and the total energy of the products is more than the total energy of the reactants.				
	A rule of thumb is that if a 20 kJ of free energy (mean which means there will es after the reaction takes pla	chemical reaction releases more hing $\Delta G < -20$ kJ), then the reaction sentially be no detectable amount ace.	e than approximately on "goes to completion," nt of the reactants left		
	Similarly, if a chemical read (meaning $\Delta G \ge +20$ kJ), the no detectable amount of p	ction would require more than 2 en "no reaction occurs," meaning products made.	0 kJ of free energy g there will essentially be		
	no reaction	some reaction	complete reaction		
	Δ <i>G</i> > +20 kJ ◀	+20 kJ $\geq \Delta G \geq -20$ kJ $\Delta G = 0$	ΔG < −20 kJ ►		
		spontaneous spontaneous			

Big Ideas	Details	Unit: Thermochemistry (Heat)
	Law	s of Thermodynamics
	Thermodynamics allows us to on energy changes. These pr thermodynamics:	predict much of what happens in the universe based edictions can be summed up in the four [*] laws of
	 If you allow objects/sy an infinite amount of ("You have to play the 	stems in contact with each other to exchange heat for time, they will have the same temperature. game.")
	 Chemical reactions pr energy (enthalpy) mo with lower temperatu to a higher temperatu ("You can't win.") 	oceed in a way that releases energy, <i>i.e.</i> , usable heat ves from particles with higher temperature to particles re. Heat energy cannot flow from a lower temperature re.
	 Some heat energy is a recovered. Therefore ("You can't break eve 	lways lost to the surroundings (entropy) and cannot be , the entropy of the universe is always increasing. n.")
	 In a completely closed constant. A change the would normally not he there is a sufficient in ("You can't get out of 	system, the total energy (enthalpy plus entropy) is at would cause an increase in enthalpy (positive ΔH) appen spontaneously, but it can occur spontaneously if crease in entropy (positive ΔS). the game.")
	Wh	at You Need to Know
	This is a lot of information, a college courses on thermody included to explain the bigge these equations. The import chemistry class are:	nd yet only the tip of the iceberg; there are entire namics. Most of the information in this section is r picture; in this course, you do not need to use any of ant points you need to understand in a first-year
	 Some energy, called er taken up or given off ir 	thalpy (<i>H</i>), exists in the form of usable heat that can be chemical reactions, and/or used to do work.
	 Some energy, called er ("lost") to the surround 	tropy (S), exists in the form of heat that is dispersed lings and cannot be recovered.
	 The combination of en determines whether an 	halpy and entropy, called Gibbs free energy (G), ad to what extent a chemical reaction occurs.
	* Originally there were three lav numbered zero in order to pre	—
	Use this space for summary a	nd/or additional notes:

Enthalpy of Formation

Unit: Thermochemistry (Heat)

MA Curriculum Frameworks (2016): HS-PS1-4

Mastery Objective(s): (Students will be able to...)

- Determine the enthalpy of formation for selected compounds by looking up data in a table.
- Identify the formation of compounds as spontaneous or non-spontaneous based on the sign of the enthalpy of formation.

Success Criteria:

- Enthalpy of formation has the correct sign and the correct units.
- Compounds are correctly identified as forming spontaneously and not forming spontaneously based on the sign of their enthalpies of formation.

Tier 2 Vocabulary: formation, spontaneous

Language Objectives:

- Explain how to find the enthalpy of formation of a compound.
- Explain how to determine whether formation of a compound is spontaneous or non-spontaneous.

Notes:

Although free energy is the best predictor of whether a reaction happens, focusing on just the changes in enthalpy can be more practical for several reasons.

- 1. Entropy cannot be measured, but must be calculated based on other measurements and calculations, such as enthalpy and equilibrium. For this reason, entropy numbers are difficult to obtain and are often not available for reactions of interest.
- 2. Free energy depends on the temperature. (Recall that G = H TS) If a reaction is producing or consuming heat, the temperature will be changing during the reaction, which means free energy calculations will require careful measurements and complex, calculus-based equations. Enthalpy does not depend on temperature, which means enthalpy numbers can be used directly regardless of the temperature at which they were measured.
- 3. Enthalpy measurements can be used directly to calculate the thermal energy (heat) produced or consumed by a reaction.

Enthalpy of Formation

Big Ideas	Details Unit: Thermochemistry (Heat)
	The concept of enthalpy of formation is that if we define the enthalpy content of a pure element to be zero [*] , then:
	 The enthalpy content of a compound can be measured by the reactions that produce the compound.
	• The enthalpy content of a compound equals the heat energy released by forming its chemical bonds, and therefore also equals the heat energy needed to break those bonds.
	standard enthalpy of formation: (ΔH_f^o) the amount of enthalpy (recoverable/usable heat) that is released when a compound is formed.
	In the "Intermolecular Forces" section starting on page 313, we saw that it takes energy to break intramolecular bonds (bonds between one molecule and another), and forming those bonds releases energy. The same is true for intermolecular bonds (bonds within a molecule).
	If a chemical compound forms spontaneously, it forms because the compound has less energy than the elements that it is formed from. When the compound is formed, the excess energy is given off as heat. Once that energy is released, there is no longer enough energy for the compound to spontaneously disintegrate, unless heat is added. These compounds have a negative enthalpy of formation.
	In some cases, energy is required to form a chemical compound. In these cases, the compound is unstable, and when a small amount of energy is added to break the bonds, a larger amount of energy is released. These compounds have a positive enthalpy of formation.
	* Recall that enthalpy includes internal energy, which includes the energy of chemical and nuclear bonds, internal induced electric or magnetic dipole moment, stress-strain energy due to deformation of solids, <i>etc.</i> , as well as energy due to motion of the molecules, atoms, nuclei, electrons, <i>etc.</i> Saying that the enthalpy is zero does not mean we are ignoring these factors—remember that enthalpy can be negative. We are simply choosing the "zero point" to be the total of those energies in a pure element.

Use this space for summary and/or additional notes:

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	Entrapy of ronnation Page. 401
Big Ideas	Details Unit: Thermochemistry (Heat)
	Determining Enthalpy of Formation
	The enthalpy of formation is defined as the energy that would need to be <i>added</i> in order to form a compound or ion directly from its elements.
	This means that pure elements in their natural state by definition have an enthalpy of formation of zero, because it takes no energy to form an element from its elements. This is true even for elements that are polyatomic in their natural state, such as N_2 , O_2 , Cl_2 , F_2 , Br_2 , l_2 , P_4 and S_8 ; even though these elements contain one or more chemical bonds, their enthalpy of formation is still defined to be zero.
	If a compound or ion forms spontaneously, the process must <i>release</i> energy. This means that compounds or ions that form spontaneously have <i>negative</i> enthalpies of formation. (Adding negative energy is mathematically the same as releasing energy.)
	Enthalpies of formation (ΔH_f^o) (and also entropies of formation) of selected compounds and elements are listed in "Table BB. Thermodynamic Data" in your Chemistry Reference Tables, on page 519.
	Examples
	The standard enthalpy of formation of CaCl ₂ is $-795.8 \frac{kJ}{mol}$. This means we would need to add $-798.2 kJ$ to make one mole of CaCl ₂ from elemental Ca and Cl ₂ .
	Adding a negative amount of heat means the reaction actually releases 795.8 $\frac{kJ}{mol}$ of heat. (Note that 795 kJ, or 795 800 J, is a lot of energy.)
	The standard enthalpy of formation of C_2H_2 (acetylene) is +226.7 $\frac{kJ}{mol}$. This means that acetylene does not form spontaneously, and we need to add 226.7 kJ of heat energy to produce one mole of acetylene gas. That energy will be released as heat when acetylene combusts to form CO_2 and H_2O .
	Use this space for summary and/or additional notes:

Enthalpy of Formation

Big Ideas	Details	Unit: Thermochemistry (Heat)	
	Homework Problem		
	Based on enthalpy of fo	prmation ($\Delta H_f^{ m o}$) data in "Table BB. Thermodynamic Data" on	
	page 519 of your Chem order, from least stable	istry Reference Tables, rank the following ten compounds in e to most stable.	
	AgCl, Al	₂ O ₃ , C ₂ H ₂ , H ₂ O, CO ₂ , N ₂ O, CuSO ₄ , H ₂ , MgSO ₄ , Si	
	1		
	2		
	3		
	4		
	5		
	6		
	7		
	8		
	9		
	10		

		Page:
Big Ideas	Details Unit: Th	ermochemistry (H
	Heat of Reaction	
	Unit: Thermochemistry (Heat)	
	Mastery Objective(s): (Students will be able to)	
	 Identify chemical reactions as exothermic or endotherm 	ic.
	 Calculate the heat produced or consumed by a reaction formation. 	based on heats of
	 Use the heat of reaction correctly in stoichiometry calcu 	lations.
	Success Criteria:	
	Heat of reaction accounts for heats of formation of all re	eactants and produ

of reaction accounts for heats of formation of all reactants and products.

- Calculation uses the correct signs for heats of formation of reactants and products, based on whether the compounds are being consumed or formed.
- Heat of formation is treated like a coëfficient (with its correct sign and units) in stoichiometry calculations.

Tier 2 Vocabulary: reaction, product

Language Objectives:

- Explain how heat of reaction is calculated.
- Explain how heat of reaction is used in stoichiometry calculations.

Notes:

Because a chemical reaction is the conversion of reactants into products, we can think of a chemical reaction as "taking apart" the reactants and then using the atoms to "build" the products.

"Taking apart" the reactants would involve breaking the bonds in compounds to form the elements they were made from. Because heats of formation give the energies to form the compounds, the energy needed to take apart the compounds would have the opposite sign. For example, if the heat of formation of H₂O is $-285.8 \frac{kJ}{mol}$, this indicates that 285.8 kJ of energy is released when one mole of H₂O is formed, which means we have to supply 285.8 kJ of energy to turn one mole of H₂O into hydrogen and oxygen.

Yes, this is confusing. Heat of formation is the amount of energy we have to put in to form the compound. A negative number means that the energy *comes out* when we form the compound instead of putting it in. But then, if we want to break apart the compound and turn it back into its elements, we have to put in that same amount of energy.

Heat of Reaction

Big Ideas	Details Unit: Thermochemistry (Heat)	
	Sample Problem:	
	Q: Calculate the energy produced by the reaction:	
	$Ca + FeO \rightarrow CaO + Fe$	
	A: First we look up our heats of formation. We need to remember to change t sign of any compounds that we are breaking apart (reactants), and keep the the same for compounds that we are creating (products):	
	Ca + FeO \rightarrow CaO + Fe $0\frac{kJ}{mol}$ +272 $\frac{kJ}{mol}$ -634.92 $\frac{kJ}{mol}$ 0 $\frac{kJ}{mol}$	
	sign changed because we're "taking apart" FeO	
	To find the heat of reaction, we simply add up all of the heat of formation numbers for each compound in the reaction:	
	$(+272 \frac{kJ}{mol}) + (-634.92 \frac{kJ}{mol}) = -361 \frac{kJ}{mol}$	
	Because the coëfficients are all one, that means we can cancel moles to give:	
	(+272 kJ) + (-634.92 kJ) = -361 kJ	
	The last step is to change the sign of the answer. Our calculation tells us that the enthalpy change for the reaction is −361 kJ. That means the products have 361 kJ <i>less</i> enthalpy than the reactants did. That energy was <i>released</i> as heat, which means the heat was <i>produced</i> by the reaction.	
	We always represent heat as a positive quantity in chemical reactions. If heat is produced, it is a product so it goes on the right. This gives us:	
	$Ca + FeO \rightarrow CaO + Fe + 361 kJ$	

Heat of Reaction

Big Ideas	Details Unit: Thermochemistry (Heat
	If we had gotten a positive number for the heat of reaction, that would mean that the products have more enthalpy content than the reactants, which means we would have had to add heat in order to make the reaction happen. In this case, heat would be a reactant and would go on the left.
	For example, in the reaction:
	$N_2 + O_2 \rightarrow 2 \text{ NO}$
	We see that N ₂ and O ₂ are elements with an enthalpy of formation of zero, and NO has an anthology of formation of 100.2 k^{1} . Because the assiftision for NO
	is 2, this means the enthalpy of reaction is +180.4 kJ. (Using heat of reaction with stoichiometry will be explained on the following pages.) The products have more enthalpy than the reactants, which means we have to add energy to make the reaction happen. This means heat is a reactant, so it goes on the left:
	N_2 + O_2 + 180.4 kJ \rightarrow 2 NO

Big Ideas	Details Unit: Thermochemistry (Heat)		
	Heat of Reaction and Stoichiometry		
	Heat of reaction works just like the coëfficients in a stoichiometry problem.		
	Because heats of reaction are expressed in $\frac{kJ}{mol}$, we need to multiply each		
	substance's heat of formation by its coëfficient in order to calculate its contribution to the heat of reaction.		
	Sample Problem:		
	Q: Calculate the heat produced by the reaction:		
	$3 \text{ Ca} + 2 \text{ AlCl}_3 \rightarrow 3 \text{ CaCl}_2 + 2 \text{ Al}$		
	A: Again we look up our heats of formation. However, now, we need to <i>both</i> remember to use positive <i>vs.</i> negative numbers correctly <i>and</i> multiply each heat of formation by its coëfficient in the chemical equation:		
	$\begin{array}{rrrrrrrrrrrrrrrrrrrrrrrrrrrrrrrrrrrr$		
	As before, to find the heat of reaction, we add the heats of formation:		
	(+1408kJ) + (-2387.4kJ) = -979kJ		
	The result tells us that the reaction produces 979 kJ of heat when the number of moles of each reactant and product are the same as the coëfficients. (This is sometimes described as "979 kJ per mole of reaction".) We can write the equation as follows:		
	3 Ca + 2 AlCl ₃ → 3 CaCl ₂ + 2 Al + 979 kJ		
	Again, notice that we changed the sign of the heat when we wrote it in the reaction because heat is produced, which means (positive) heat is a product of the reaction.		
	In this reaction, suppose we had used 4.5 mol of Ca (and excess AlCl ₃). To find the heat produced, we would calculate it as if the 979 kJ were the coëfficient for heat in the chemical equation.		
	Using our stoichiometry "input/output machine", 3 mol of Ca produces 979 kJ of heat. This means that to get from 3 to 979, we have to divide by 3 and multiply by 979. This gives:		
	$\frac{4.5 \text{ motCa}}{1} \cdot \frac{979 \text{ kJ}}{3 \text{ motCa}} = 1469 \text{ kJ of heat produced}$		
	Use this space for summary and/or additional notes:		

Heat of Reaction

Big Ideas	Details	Unit: Thermochemistry (Heat)	
	Homework Problems		
	Using enthalpy of formation (ΔH Chemistry Reference Tables, on the following chemical reactions correct place.	f_{f}^{o}) data in "Table BB. Thermodynamic Data" in your page 519, calculate the heat of reaction for each of and write out the reaction with the heat in the	
	1. Pb + FeSO ₄ \rightarrow PbSO ₄ + Fe		
	Answer: $\Delta H^o = +9 \text{kJ}$ Pb	+ FeSO ₄ + 9 kJ \rightarrow PbSO ₄ + Fe	
	2. 2 C_2H_2 + 5 $O_2 \rightarrow$ 4 CO_2 + 2	H ₂ O	
	Answer: $\Delta H^{\circ} = -2511 \text{ k}$	$J = 2 C_2 H_2 + 5 O_2 \rightarrow 4 CO_2 + 2 H_2 O + 2511 kJ$	
	3. $C_6H_{12}O_6 + 6 O_2 \rightarrow 6 CO_2 +$	6 H ₂ O	
	Answer: $\Delta H^{\circ} = -2537 \mathrm{k}$	J $C_6H_{12}O_6 + 6 O_2 \rightarrow 6 CO_2 + 6 H_2O + 2 537 kJ$	

Bond Energies

Unit: Thermochemistry (Heat)

MA Curriculum Frameworks (2016): HS-PS1-4

Mastery Objective(s): (Students will be able to...)

- Identify the bonds in molecular compounds and look up their bond energies.
- Calculate the heat produced or consumed by a reaction based on bond energy calculations.

Success Criteria:

Details

Big Ideas

- Bond energies have the correct sign and the correct units based on where they appear in the chemical equation.
- Heat of reaction calculated from bond energies agrees with heat of reaction calculated from heat of formation data.

Tier 2 Vocabulary: bond

Language Objectives:

• Explain how heat of reaction is calculated using bond energies.

Notes:

lonic compounds form crystals that have straightforward ratios of cations to anions, which means it is practical to publish a table of the heats of formation of common ionic compounds and to calculate the heat of reaction from heat of formation data. Molecular compounds, however, can form long chains, and there are thousands upon thousands of common molecular compounds.

For most molecular compounds, instead of attempting to publish an exhaustive list of heat of formation data, it is much simpler to publish a table of the energy it takes to dissociate the individual bonds in molecular compounds. This way, it is possible to estimate the heat of reaction by simply adding up the energies for breaking or forming the individual bonds in a compound.

These bond energies are shown in "Table AA. Bond Dissociation Energies & Bond Lengths" in your Chemistry Reference Tables on page 518. Note that the table is "bond *dissociation* energies." This means that the numbers represent the amount of energy needed to *dissociate (break)* the bonds listed. If it takes energy (a positive number) to break a bond, that means energy is released (a negative number) when the bond is formed. This means we can use bond energies to calculate the enthalpy of formation of a compound, but we need to remember that the bond energies are negative when bonds are formed.
Bond Energies

Here are bond dissociation energies for some bonds commonly found in organic compounds. A more extensive list can be found in "Table AA. Bond Dissociation Energies & Bond Lengths" on page 518 of your Chemistry Reference Tables.

Bond	C-C	C=C	C≡C	C-H	C-0	C=0	0=0	0-н
Bond Dissociation Energy [*] (^{kJ} _{mol})	346	602	835	411	358	799	494	459

Sample problem:

Details

Big Ideas

Use bond dissociation energy data to predict the energy of reaction for the combustion of methane:

$$\mathsf{CH}_4 + 2 \ \mathsf{O}_2 \rightarrow \ \mathsf{CO}_2 + 2 \ \mathsf{H}_2\mathsf{O}_2$$

This actually means:

н—С—н Н	+	2	0=0	\rightarrow	0=C=0	+	2	HH
break 4 C–H bonds	+	2	break 1 O=O bond	\rightarrow	form 2 C=O bonds	+	2	form 2 O–H bonds
4(411 kJ)	+	2	(494 kJ)	\rightarrow	2(-799 kJ)	+		4(–459 kJ)
1 644 kJ	+		988 kJ	+	-1 598 kJ	+		-1 836 kJ

Note that numbers on the left are positive, because we are *dissociating (breaking)* those bonds, and numbers on the right are negative because we are *forming* those bonds.

As with heat of reaction calculations, we add the energies for each of the individual molecules:

^{*} These are bond dissociation energies are for homolytic dissociation (*i.e.*, the electrons are equally split between the two atoms). Heterolytic bond dissociation energies (where the electrons split unequally) are always higher.

	Bond Energies Page: 4						
Big Ideas	Details Unit: Thermochemistry (Heat						
	1 644 kJ + 988 kJ + (-1 598 kJ) + (-1 836 kJ) = -802 kJ						
	We can now write the reaction as:						
	($CH_4 + 2 O_2 \rightarrow CO_2$	2 + 2 H ₂ O + 802 k	L			
	This agrees with the heat of reaction that we would calculate from enthalpy of formation data:						
		Compound	$\Delta H_f^\circ \left(rac{kJ}{mol} ight)$				
		CH ₄ (g)	-74.8				
		O ₂ (g)	0				
		CO ₂ (g)	-393.5				
		$H_2O(g)$	-241.8				
	CH4 +	$2 O_2 \rightarrow$	CO ₂ +	2 H ₂ O			
	+74.8 ^{kJ} mol	2(0 <u>kJ</u> mol)	–393.5 ^{_kj} mol	2(–241.8 ^{kJ} _{mol})			
	+74.8 kJ	0 kJ	–393.5 kJ	–483.6 kJ			
	74.8 kJ -	+ 0 + (-393.5 kJ)	+ (-483.6 kJ) = -	802.3 kJ			
	Note that adding the bond compound's enthalpy of for	d dissociation en ormation. As an	ergies for a com example, recall	pound does not give the that the heat of formation			
	of O ₂ is zero, but it takes 4	$494 \frac{kJ}{mol}$ of heat e	energy to break a	an O=O double bond.			

Bond Energies

		Dona L	ICI BIC				Fage. 471
Big Ideas	Details				Uni	t: Therm	ochemistry (Heat)
	Limitations of Using Bond Energy to Calculate Heat of Reaction						
	The above example with methane is one of the few instances where bond energies and heats of formation give the same values. This is often not the case. For example, if you calculate the heat of reaction for the combustion of acetylene:						
		2 H−CΞC−+	+ 5 02	\rightarrow 4 CO	2 + 2 H ₂ C) + heat	
	you would get −2 511 kJ using heat of formation data, but −4 914 kJ using bond energies.						
	Enthalpies of formation are well-defined and precise, meaning that different people could measure them and expect to get the same answer. Average bond enthalpies, however, can vary widely depending on which molecules were used to gather the data. (In fact, the numbers may vary significantly from one published table to another.)						
	For example, there of a C–H bond in C acetylene.	e are many diffe CH₄ is very differe	ent mol ent from	lecules v the boi	with C–H nd energ	I bonds. gy of a C-	The bond energy –H bond in
	The bond energies combustion of me	found in the bo thane are based	nd ener on:	gy table	s for the	e bonds i	involved in the
		Bond	C-H	C=O	0=0	O-H]
		Energy $\left(\frac{kJ}{mol}\right)$	411	799	494	459	
		Based on	CH_4	CO ₂	O ₂	H ₂ O	
	Needless to say, th bonds broken or fo surprising that usin more energy to br using bond energin Because average b formation, calcular quick estimation, o	ne basis for each ormed in the cor ng bond energies eak the C–H bon es does not give bond enthalpies a ting heat of reac or as a last resor	of these nbustion s gives t d in ace the corr are muc tion fron t when h	e bond e n of met he same tylene, rect answ h less ac m bond neat of f	energies thane. If e result. which is wer for a ccurate t energies formatio	is exacti t is there part of f acetylen than ent s should n data a	ly the same as the efore not er, it takes much the reason that e. halpies of only be used for a are not available.

Bond Energies

Big Ideas	Details	Unit: Thermochemistry (Heat)
		Homework Problem
	1.	Propane (C ₃ H ₈) has the following structure:
		Н Н Н H—С—С—С—Н Н Н Н
		Using the bond energy data from this chapter or on page 518 of your Chemistry reference tables, compute the heat of reaction for the complete combustion of propane:
		$C_3H_8 + 5 O_2 \rightarrow 3 CO_2 + 4 H_2O + heat$

Answer: –2 016 kJ (Note: calculating from ΔH_f^o values gives 2 044 kJ)

Dia Ideaa	Detaile	1/5					
Big ideas	Details Unit: Kinetics & Equilibri	um					
	Introduction: Kinetics & Equilibrium						
	Unit: Kinetics & Equilibrium						
	Topics covered in this chapter:						
	Collision Theory47	4					
	Rate of Reaction (Kinetics)47	7					
	Equilibrium48	1					
	Le Châtelier's Principle48	8					
	Standards addressed in this chapter:						
	Massachusetts Curriculum Frameworks & Science Practices (2016):						
	HS-PS1-4 Develop a model to illustrate the energy transferred during an exothermic or endothermic chemical reaction based on the bond energy difference between bonds broken (absorption of energy) and bonds form (release of energy).	ed					
	HS-PS1-5 Construct an explanation based on kinetic molecular theory for why varying conditions influence the rate of a chemical reaction or a dissolving process. Design and test ways to slow down or accelerate rates of process (chemical reactions or dissolving) by altering various conditions.	/ ; ;es					
	HS-PS1-6 Design ways to control the extent of a reaction at equilibrium (relat amount of products to reactants) by altering various conditions using Le Châtelier's principle. Make arguments based on kinetic molecular theory account for how altering conditions would affect the forward and reverse rates of the reaction until a new equilibrium is established.	ive :o					

Unit: Kinetics & Equilibrium

MA Curriculum Frameworks (2016): HS-PS1-4, HS-PS1-5

Mastery Objective(s): (Students will be able to...)

• Use collision theory to explain activation energy and when a reaction does or does not occur.

Success Criteria:

Details

- Explanations refer to Kinetic Molecular Theory (KMT).
- Explanations account for the role of activation energy.

Tier 2 Vocabulary: collision

Language Objectives:

• Explain the process of reactants colliding with enough energy to react and form products.

Notes:

Recall that Kinetic Molecular Theory (KMT) states:

- Gases are made of very large numbers of molecules.
- Molecules are constantly moving (obeying Newton's laws of motion), and their speeds are constant.
- Molecules are very far apart compared with their diameter.
- Molecules collide with each other and walls of container in elastic collisions.
- Molecules behaving according to KMT are not reacting or exerting any other forces (attractive or repulsive) on each other.

Collision Theory

Big Ideas Details Unit: Kinetics & Equilibrium When we studied KMT earlier this year, we mentioned that the last bullet point did not mean that reactions did not occur, but that they were covered by collision theory, not KMT. Collision theory states that: Molecules can be modeled as rigid spheres. Molecules move according to KMT. • When molecules collide, kinetic energy (proportional to temperature) is transferred. • If the molecules collide with enough energy to break existing chemical bonds, and if they are oriented in a way that allows the new bonds to form, a reaction occurs. • If a reaction does not occur, the reactants "bounce" off each other as described by KMT. For example, consider the reaction: $NO_3 + CO \rightarrow NO_2 + CO_2$ Reactants moving Ineffective collision too slowly (molecules bounce) Reactants not facing Ineffective collision the right direction (molecules bounce) Reactants moving Effective collision fast enough and (reaction happens) facing the right direction effective collision: a collision with enough energy that it leads to formation of products ineffective collision: a collision that does not have enough energy to lead to formation of products. mechanism: the details of which molecules collide and in what order and orientation in order for the reaction to take place.



Unit: Kinetics & Equilibrium

MA Curriculum Frameworks (2016): HS-PS1-5

Mastery Objective(s): (Students will be able to...)

• List and explain factors that affect the rate of a chemical reaction.

Success Criteria:

Details

• Descriptions convey how each factor affects the rate of reaction.

Tier 2 Vocabulary: intermediate

Language Objectives:

• Explain what it means for a reaction to happen faster *vs.* slower, and how each factor affects the reaction rate.

Labs, Activities & Demonstrations:

• Drop of food coloring in hot vs. cold water.

Notes:

- <u>reactants</u>: the compounds consumed in the chemical reaction; compounds that *react*.
- <u>products</u>: the compounds created by the chemical reaction; compounds that are *produced*.
- <u>intermediates</u>: compounds that are produced in one step of a multi-step reaction and consumed by a later step.

Rate of Reaction (Kinetics)

	rage: 476
Big Ideas	Details Unit: Kinetics & Equilibrium
	<u>reaction rate</u> (<i>k</i>): the rate at which products are formed in a chemical reaction,
	usually expressed in units of: $\frac{\text{mol}}{\text{L}\cdot\text{s}}$ or $\frac{\text{M}}{\text{s}}$ (where M = molarity = $\frac{\text{mol}}{\text{L}}$)
	The reaction rate is related to the activation energy. A reaction with higher activation energy will happen more slowly, because fewer of the collisions will have enough energy to enable the molecules to react. Conversely, a reaction with lower activation energy will happen more quickly.
	The equation for rate of reaction is: $\ln(k) = -\frac{RT}{E_a}$ or $k = e^{-RT/E_a}$
	Quantitative rate calculations are studied in AP^{\circledast} Chemistry. In this course, you need to understand how the equation shows that a higher temperature will speed up the reaction (larger value of k), and a higher activation energy will slow down the reaction (smaller value of k).
	<u>rate-limiting step</u> (or <u>rate-determining step</u>): the step that determines the overall rate of the reaction. In a multi-step reaction, the rate-limiting step is the slowest step.
	For example, in the multi-step reaction:
	$A \xrightarrow{\text{fast}} B \xrightarrow{\text{slow}} C \xrightarrow{\text{fast}} D$
	 A → B will happen faster than B can get used up, so B will accumulate and the first reaction will not affect the overall rate.
	 C → D will happen fast, which means as soon as C is produced, it will react to produce D.
	Therefore, the rate of B \rightarrow C, which happens slowly, is what determines the overall rate of the reaction A \rightarrow D.
	<u>catalyst</u> : a substance that speeds up a reaction by lowering the activation energy of (and therefore speeding up) the rate-limiting (slowest) step.

Big Ideas	Details Unit: Kinetics & Equilibrium
	Factors that Affect Reaction Rates
	 <u>concentration of reactants</u>: higher concentration means more frequent collisions = faster rate. (Only applies to molecules involved in the rate- determining step.) For gases, higher pressure = higher concentration.
	 <u>surface area of reactants</u>: more surface area means higher probability of a collision = faster rate.
	• <u>temperature</u> : higher temperature = faster because faster-moving molecules collide more often, and because faster-moving molecules have more kinetic energy to overcome the activation energy.
	 <u>nature of the reactants</u>: weak bonds are easier to break than strong bonds. Reactions involving dissolved ions are very fast, because bonds are already broken.
	 <u>catalysts</u>: catalysts <u>speed up reactions</u> in any of several ways:
	 bring molecules into the correct orientation for an effective collision (equivalent to increasing the concentration and/or surface area)
	 assist in breaking of bonds in the reactant(s) and/or formation of bonds in the products (equivalent to changing the nature of the reactants and/or lowering the activation energy)
	Catalysts are not reactants; they are <u>not</u> consumed by the reaction.

Rate of Reaction (Kinetics)

Nate of Neuclion	(INITE LIES)	rage. 4
Details		Unit: Kinetics & Equilibriu
Homewo	rk Problen	ns
Consider the following decomposition re	eaction:	
2 N ₂ O ₅ -	→ 2 N ₂ + 5 O ₂	
This reaction happens in three steps:		
1. 2 N ₂ O ₅ + 2 H ₂ O \rightarrow 4 HNO ₃	fast	
2. 2 HNO ₃ \rightarrow N ₂ + 3 O ₂ + H ₂	slow	
3. $2 H_2 + O_2 \rightarrow 2 H_2 O$	very fast	
Answer the following questions:		
1. Which compounds are intermed	liates in this rea	action?
 If you wanted to speed up the o would you try to speed up? Exp might do this. 	verall reaction, lain why, and g	, which of the three steps ive an example of how you

Use this space for summary and/or additional notes:

Big Ideas

	Equilibrium Page: 48
g Ideas	Details Unit: Kinetics & Equilibriur
	Equilibrium
	Unit: Kinetics & Equilibrium
	MA Curriculum Frameworks (2016): HS-PS1-6
	Mastery Objective(s): (Students will be able to)
	 List and explain factors that affect the equilibrium of a chemical reaction.
	 Write equilibrium expressions from chemical equations and chemical equations from equilibrium expressions.
	Success Criteria:
	 Equilibrium expressions have products on top and reactants on the bottom.
	 Equilibrium expressions have coëfficients shown as exponents.
	Tier 2 Vocabulary: forward, backward, expression
	Language Objectives:
	• Explain how different factors affect the equilibrium of a chemical reaction.
	Notes:
	reversible reaction: a reaction that proceeds in both forward and backward directions. Usually written with double half-arrows:
	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
	dynamic chemical equilibrium: reaction is happening in both directions, but the changes balance each other, so the concentrations of reactants & products remain constant. At equilibrium:
	rate _{forward} reaction = rate _{reverse} reaction or $k_f = k_r$
	<u>concentration</u> : the amount of a compound dissolved in a given amount of solution,
	concentration: the amount of a compound dissolved in a given amount of solution, usually expressed in $\frac{mol}{L}$ (moles of the compound per liter of solution).

Big Ideas

Details

	Arrows Used in Chemical Equations					
	Arrow	Meaning				
	$A + B \rightarrow C + D$	A + B react to produce C + D Either there is little or no reverse reaction, or no information is given about equilibrium.				
	$A + B \rightleftharpoons C + D$	A + B are in equilibrium with C + D No information is given about whether products or reactants are favored.				
	$A + B \xrightarrow{\sim} C + D$	A + B are in equilibrium with C + D Products are favored. (<i>I.e.</i> , the concentrations of products are higher than the concentrations of reactants.)				
	A + B <u>→</u> C + D	A + B are in equilibrium with C + D Reactants are favored. (<i>I.e.,</i> the concentrations of reactants are higher than the concentrations of products.)				
-	A↔B	A and B are different resonance structures of the same compound. This is different from a chemical reaction that is at equilibrium.				

Fauilibrium

leas	Details	Unit: Kinetics & Equilibrium
	equilibrium expression: a mathematical expression retrieved the products and reactants at equilibrium. For the	elating the concentrations of ne reaction:
	$N_2 + 3H_2 \rightleftharpoons 2NH_3$	
	the equilibrium expression is:	
	[NH ₂] ²	
	$\mathcal{K}_{eq} = \frac{[N_{eq}]^3}{[N_2][H_2]^3}$	
	where $[NH_3]$ means the molarity of NH_3 (the cond	centration in $\frac{mol}{L}$).
	The equilibrium expression comes from collision to react, one N_2 and 3 H_2 molecules need to find this 4-way collision is related to the product of th molecules involved, <i>i.e.</i> , $[N_2][H_2][H_2][H_2]$ or $[N_2][$ coëfficients in the equation became exponents in	theory. In order for N_2 and H_2 each other. The probability of e concentrations of each of the H_2] ³ . Notice that the n the equilibrium expression.
	The reverse reaction requires 2 NH_3 molecules to collision is the product of the concentrations, [NH coëfficient "2" became the exponent.	collide; the probability of this H_3 [NH ₃] or [NH ₃] ² . Again, the
	For the equilibrium expression, we write the cond top and reactants on the bottom. (The reason fo for the forward and reverse reactions. Rate laws course, but are studied in AP Chemistry.) Howev means that an equilibrium constant greater than products, and an equilibrium constant less than 1 reactants.	centrations of products on the r this comes from the rate laws are beyond the scope of this er, for our purposes, it also 1 means we have more means we have more
	equilibrium constant (K_{eq}): when we plug the concen	trations (in $\frac{mol}{l}$) for each of the
	products and reactions into the equilibrium expre	ession, we get a value for K_{eq} .
	If $K_{eq} > 1$ then there are more products than react "equilibrium lies to the right".	ants, and we say that the
	If $K_{eq} < 1$ then there are more reactants than proc "equilibrium lies to the left".	ducts, and we say that
	If $K_{eq} = 1$ then there are equal amounts of reactar that the reactants and products are equal	nts and products, and we say qually favored.
	K_{eq} depends on temperature, but is constant for a temperature.	a given reaction at a given



Equilibrium

Ideas	Details Unit: Kinetics & Equilibrium
	Q: Calculate the value of the equilibrium constant for the reaction $2 \text{ NOBr} \rightleftharpoons 2 \text{ NO} + \text{Br}_2$ if the concentration of NOBr is 3.00 M, the concentration of NO is 0.750 M, and the concentration of Br ₂ is 0.200 M
	A: The equilibrium expression is:
	$K_{eq} = \frac{[\text{NO}]^2[\text{Br}_2]}{[\text{NOBr}]^2}$
	Plugging in the concentrations, we get:
	$K_{eq} = \frac{(0.750)^2 (0.200)}{3.00^2} = 0.0125$
	Note that the units for equilibrium constants are tricky, because each concentration is in $\frac{mol}{L}$. This means that if a concentration is squared, the units
	become $\left(\frac{\text{mol}}{L}\right)^2 = \frac{\text{mol}^2}{L^2}$. This means that every equilibrium constant has its own
	units, depending on the number of molecules that take part in the forward and reverse reactions.
	From the equilibrium expression, the units of the equilibrium constant for this expression happens to be:
	$\frac{\left(\frac{\text{mol}}{L}\right)^2\left(\frac{\text{mol}}{L}\right)}{\left(\frac{\text{mol}}{L}\right)^2} = \frac{\text{mol}}{L}$
	Working with the units for the equilibrium constant is beyond the scope of this course.

Equilibrium

Big Ideas	Jetails Unit	:: Kinetics & Equilibri
	Homework Problems	
	Vrite the expression for the equilibrium constants for each o eaching actions.	f the following
	1. $Xe + 3F_2 \rightleftharpoons 2XeF_6$	
	2. $CH_4 + 2H_2S \rightleftharpoons CS_2 + 4H_2$	
	3. $3 \text{CO}_2 + 4 \text{H}_2 \text{O} \rightleftharpoons \text{C}_3 \text{H}_8 + 5 \text{O}_2$	
	4. Write the chemical equation for the equilibrium syste	em given by the
	expression: $K_{eq} = \frac{[H_2O]^2[O_2]}{[H_2O_2]^2}$	
	5. Write the chemical equation for the equilibrium syste expression:	em given by the
	$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$	
	 Write the chemical equation for the equilibrium syste expression: 	em given by the
	$\kappa_{eq} = \frac{[\text{HCI}]^4[\text{O}_2]}{[\text{H}_2\text{O}]^2[\text{CI}_2]^2}$	
	Ise this space for summary and/or additional notes:	

		Equilibrium Page: 48	57
Big Ideas	Details	Unit: Kinetics & Equilibriu	m
	7.	A reaction vessel contains 0.150 M CH ₄ , 0.233 M H ₂ O, 0.259 M H ₂ , and 0.513 M CO. If the equilibrium reaction is $CH_4 + H_2O \rightleftharpoons CO + 3H_2$, write the equilibrium expression and calculate the value of K_{eq} .	š
		Answer: $K_{eq} = 0.255$	
	8.	Answer: $K_{eq} = 0.235$ A 10 L flask contains 0.128 mol of CO, 0.155 mol of H ₂ and 0.0244 mol of CH ₃ OH. If the equilibrium reaction is CH ₃ OH \rightleftharpoons CO+2H ₂ , write the equilibrium expression and calculate the value of K_{eq} . (<i>Note: you will need to divide each number of moles by 10 L to get the concentrations in</i> $\frac{mol}{L}$.)	
		Answer: $K_{eq} = 0.00126$	

Le Châtelier's Principle

Unit: Kinetics & Equilibrium

MA Curriculum Frameworks (2016): HS-PS1-6

Mastery Objective(s): (Students will be able to...)

• Use Le Châtelier's Principle to predict a shift in equilibrium in response to a change.

Success Criteria:

• Prediction correctly describes the shift in equilibrium when the concentration of one chemical species is changed.

Tier 2 Vocabulary: stress

Language Objectives:

• Explain how a change provokes a response.

Notes:

If a reaction is at equilibrium, the reaction will resist any change with a corresponding change that shifts the reaction back to its equilibrium. Because K_{eq} is a constant, after the equilibrium shifts, the value of K_{eq} will be the same as it was before the change.

In plain English, if you change something, the equilibrium will shift to partly undo the change. This principle is called Le Châtelier's Principle, named after the French chemist Henry Louis Le Châtelier who first proposed the idea.

For example, consider the reaction:

$$N_2(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g) + 92.1 kJ$$

For this reaction, $K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3} = 835$ at 25 °C.

Suppose we started with $[N_2] = 0.05 \text{ M}$, $[H_2] = 0.3 \text{ M}$, and $[NH_3] = 1.06 \text{ M}$.

If we add more $[H_2]$, the reaction would use more H_2 , and make more NH_3 . If we kept adding H_2 until $[H_2] = 0.4$ M, we would have $[N_2] = 0.026$ M, and $[NH_3] = 1.18$ M. As you can see, adding more H_2 caused the reaction to use up more N_2 and make more NH_3 .

Le Châtelier's Principle

Details Unit: Kinetics & Equili Le Châtelier's Principle tells us that we don't have to perform the equilibrium calculation to qualitatively predict what will happen. We can just look at the equation:

N_2 (g) + 3 H_2 (g) \Rightarrow 2 NH_3 (g) + 92.1 kJ

if we add more H_2 , the equilibrium will shift to use more of it up. This means the equilibrium will shift to the right, also using up more N_2 and making more NH_3 .

On the other hand, if we added NH_3 , the equilibrium would instead shift to the left to use up some of the NH_3 , and make more N_2 and H_2 .

Action	Equilibrium shift
Add N_2 or H_2	to the right
Remove N_2 or H_2	to the left
Add NH ₃	to the left
Remove NH ₃	to the right
Increase the temperature (add heat)	to the left

Note that the value of K_{eq} is different at different temperatures. Adding reactants or products doesn't change the value of K_{eq} , but changing the temperature does. Le Châtelier tells us that adding heat must shift the equilibrium to the *left*. The equilibrium shift occurs because increasing the temperature results in a lower value of K_{eq} for this equation.

Quantitative equilibrium calculations and the relationship between the equilibrium constant and thermodynamics are studied in more depth in AP[®] Chemistry.

Use this space for summary and/or additional notes:

Big Ideas

Le Châtelier's Principle

Big Ideas	Details		Unit: Kinetics & Equilibrium	
		Homework Problem	S	
	Consider the chemical equ	uation:		
		$6 H_2 (g) + P_4 (g) \rightleftharpoons 4 PH_3 (g) + 53.5$	kJ	
	 Indicate which dire following: 	 Indicate which direction the equilibrium would shift as a result of each of following: 		
	a. Adding P_4	c. Rem	noving H_2	
	b. Removing PH₃	d. Dec	reasing the temperature	
	2. Write the equilibri	um expression for the above re	eaction.	
	3. The value of K _{eq} for equilibrium at 25 ° of P₄ is 0.025 M, w	r this reaction is 4.44 at 25 °C. C, the concentration of H_2 is 1.0 hat is the concentration of PH_3	If the reaction is at 00 M and the concentration ?	
	Answer: [PH₃] = 0. 4. If the reaction is co increases to 4.77. Châtelier's Principl	58 M poled to 4 °C, the value of the e Is this consistent with the pred e in question #1d above? Expla	quilibrium constant iction made by Le ain.	
	Use this space for summa	ry and/or additional notes:		

Details **Big Ideas Introduction: Acids & Bases** Unit: Acids & Bases Topics covered in this chapter: pH & Indicators......497 p*K*_a & Buffers......504 Standards addressed in this chapter: Massachusetts Curriculum Frameworks & Science Practices (2016): HS-PS1-9(MA) Relate the strength of an aqueous acidic or basic solution to the extent of an acid or base reacting with water as measured by the hydronium ion concentration (pH) of the solution. Make arguments about the relative strengths of two acids or bases with similar structure and composition.

Acids & Bases

Unit: Acids & Bases

Details

MA Curriculum Frameworks (2016): HS-PS1-9(MA)

Mastery Objective(s): (Students will be able to...)

- Define acids and bases based on both the Arrhenius and Brønsted-Lowry theories and give examples.
- Classify acids and bases as strong or weak.
- Identify conjugate acid and base pairs.

Success Criteria:

• Prediction correctly describes the shift in equilibrium when the concentration of one chemical species is changed.

Tier 2 Vocabulary: stress

Language Objectives:

• Explain how a change provokes a response.

Notes:

Acids are one of the first substances that come to mind when we think of chemistry. Acids are the dangerous chemicals that mad scientists in movies throw at people, and the chemicals that impressively dissolve metals and other substances right before your eyes.

Acids have held this sort of fascination for centuries. The American chemist Ira Remsen wrote the following wonderful anecdote of his first encounter with them:

While reading a text book of chemistry, I came upon the statement, "nitric acid acts upon copper." I was getting tired of reading such absurd stuff and I determined to see what this meant. Copper was more or less familiar to me, for copper cents were then in use. I had seen a bottle marked "nitric acid" on a table in the doctor's office where I was then "doing time!" I did not know its peculiarities, but I was getting on and likely to learn. The spirit of adventure was upon me. Having nitric acid and copper, I had only to learn what the words "act upon" meant. Then the statement "nitric acid acts upon copper," would be something more than mere words.

Acids & Bases

Big Ideas	Details	Unit: Acids & Bases
	All was still. In the interest of knowledge I was even willing to sacri the few copper cents then in my possession. I put one of them on opened the bottle marked "nitric acid;" poured some of the liquid copper; and prepared to make an observation. But what was this w thing which I beheld? The cent was already changed, and it was no change either. A greenish blue liquid foamed and fumed over the o over the table. The air in the neighborhood of the performance be colored dark red. A great cloud arose: This was disagreeable and s how should I stop this?	fice one of the table; on the vonderful small cent and came suffocating—
	I tried to get rid of the objectionable mess by picking it up and thro the window, which I had meanwhile opened. I learned another fac not only acts upon copper but it acts upon fingers. The pain led to unpremeditated experiment. I drew my fingers across my trousers fact was discovered. Nitric acid acts upon trousers.	wing it out t—nitric acid another and another
	Taking everything into consideration, that was the most impressive experiment, and, relatively, probably the most costly experiment I performed. I tell of it even now with interest. It was a revelation to resulted in a desire on my part to learn more about that remarkabl action. Plainly the only way to learn about it was to see its results, experiment, to work in a laboratory.	have ever o me. It e kind of to
	However, not all acids are this dangerous, especially when they a eat and drink vinegar (dilute acetic acid), orange juice (which con and Coca-Cola (which contains phosphoric acid).	re more dilute. We tains citric acid),
	acid: a substance that can produce H ₃ O ⁺ ions in water, release H and/or accept electrons from another substance.	$^{+}$ ions * in solution,
	base: the "opposite" of an acid; a substance that can produce OF accept H ⁺ ions in solution and/or donate electrons to anothe	H [−] ions in water, r substance.
	* Note that an H ⁺ ion is a proton. Chemists often use the term "proton" for convenience. Thus an acid is a compound that releases protons in protons are just the H ⁺ ions—there's no nuclear weirdness going on!	in place of "H ⁺ ion" water. These

Details Some Properties Acids • taste sour • react with some metals (<i>i.e.,</i> the ones above hydrogen on the activity series) • dissolve plants Bases • taste bitter	Unit: Acids & Bases of Acids & Bases • produce H ₃ O ⁺ ions in water (Arrhenius definition) • release H ⁺ ions (Brønsted-Lowry definition) • accept electrons (Lewis definition)
Some Properties Acids • taste sour • react with some metals (<i>i.e.</i> , the ones above hydrogen on the activity series) • dissolve plants Bases • taste bitter	 of Acids & Bases produce H₃O⁺ ions in water (Arrhenius definition) release H⁺ ions (Brønsted-Lowry definition) accept electrons (Lewis definition)
 Acids taste sour react with some metals (<i>i.e.</i>, the ones above hydrogen on the activity series) dissolve plants Bases taste bitter 	 produce H₃O⁺ ions in water (Arrhenius definition) release H⁺ ions (Brønsted-Lowry definition) accept electrons (Lewis definition)
 taste sour react with some metals (<i>i.e.</i>, the ones above hydrogen on the activity series) dissolve plants Bases taste bitter 	 produce H₃O⁺ ions in water (Arrhenius definition) release H⁺ ions (Brønsted-Lowry definition) accept electrons (Lewis definition)
• taste bitter	
• taste bitter	
 feel "slippery" (like soap) dissolve people (skin) 	 produce OH⁻ ions in water (Arrhenius definition) accept H⁺ ions (Brønsted-Lowry definition) give electrons (Lewis definition)
dissociation: to dissolve by splitting into p	ositive and negative ions
Acids & bases dissociate in water.	
Strong acids & bases dissociate compl partially.	etely; weak acids & bases only dissociate
For example: HCl - NaOH -	→ $H^+ + CI^-$ → $Na^+ + OH^-$
neutralization: a reaction in which an acic of ionic compound) plus water. For ex	l and a base react to produce a salt (a type cample:
HNO ₃ + KOH (acid) + (base)	$\rightarrow \text{KNO}_3 + \text{H}_2\text{O}$ $\rightarrow \text{(salt)} + \text{(water)}$
<u> </u>	dissociation: to dissolve by splitting into p Acids & bases dissociate in water. Strong acids & bases dissociate compl partially. For example: HCl - NaOH - <u>neutralization</u> : a reaction in which an acid of ionic compound) plus water. For ex HNO ₃ + KOH (acid) + (base)

Acids & Bases

Big Ideas	Details Unit: Acids & Bases
	<u>strong acid</u> : an acid that dissociates completely in water and produces H^+ ions, which then convert H_2O molecules to H_3O^+ ions. Strong acids include HCl, HBr, HI, H_2SO_4 , HNO ₃ , and HClO ₄ .
	weak acid: an acid that only partially dissociates in water. HF is an example of a weak acid:
	strong base: a base that dissociates completely in water and produces OH ⁻ ions. Strong bases include all of the group 1 hydroxides (LiOH, NaOH, KOH, etc.), plus the group 2 hydroxides Ca(OH) ₂ , Sr(OH) ₂ , and Ba(OH) ₂ .
	weak base: a base that only partially dissociates in water. NH ₃ is an example of a weak base:
	$NH_3 + H_2O \stackrel{\sim}{\leftarrow} NH_4^+ + OH^-$
	<u>conjugates</u> : the acid & base forms of a compound. The acid form has an extra H ⁺ that can dissociate. The base form is the same compound without the H ⁺ .
	<u>conjugate base</u> : the base formed by removing H ⁺ from an acid. For example, the conjugate base of HCl is Cl ⁻ .
	conjugate acid: the acid formed by adding H^+ to a base. For example, the conjugate acid of NH_3 is NH_4^+ .
	polyprotic: an acid that can lose more than one H ⁺ . For example, H ₂ SO ₄ can lose one H ⁺ to dissociate into H ⁺ and HSO ₄ ⁻ . HSO ₄ ⁻ can then lose a second H ⁺ to dissociate into H ⁺ and SO ₄ ²⁻ .
	Remember that an H^+ ion is just a proton. A polyprotic acid is just an acid with more than one proton that it can lose by dissociation.
	<u>amphoteric</u> : a substance that "can go either way"— <i>i.e.,</i> it has both a conjugate acid and a conjugate base. For example, the HSO_4^- ion is amphoteric:
	$H_2SO_4 \longrightarrow HSO_4^- \longrightarrow SO_4^{2-}$

Big Ideas	Details				Unit: A
		Homework Pro	bl	ems	
	Give th	e conjugate base for each of the followin	g ac	cids:	
	1.	HCI	5.	HSO_4^-	
	2.	H ₂ S	6.	H_2SO_3	
	3.	HCO3 [−]	7.	NH ₃	
	4.	H₂PO₄ [−]	8.	HS⁻	
	Give th	e conjugate acid for each of the following	g ba	ses:	
	9.	HSO₄ [−]	, 14.	PO4 ³⁻	
	10.	SO ₃ ²⁻	15.	CH₃NH₂	
	11.	ClO ₄ -	16.	F⁻	
	12.	H ₂ PO ₄ ⁻			
	13.	SO₄²-			
	10.				

Big Ideas	Details Unit: Acids & Bases
	pH & Indicators
	Unit: Acids & Bases
	MA Curriculum Frameworks (2016): HS-PS1-9(MA)
	Mastery Objective(s): (Students will be able to)
	 Calculate pH from [H⁺] and pOH from [OH⁻].
	 Identify acids and bases from their pK_a values.
	 Select an appropriate indicator for a desired pH range.
	Success Criteria:
	 pH and pOH are calculated correctly.
	 Acids and bases are correctly identified from their pK_a values.
	 Indicator changes color in a pH range that includes the pH of the given acid or base.
	Tier 2 Vocabulary: acid, base, indicator
	Language Objectives:
	 Explain why higher [H⁺] results in a lower pH.
	Notes:
	In water, a very small amount of H_2O dissociates into H^+ and OH^- ions:
	$H_2O \Longrightarrow H^+ + OH^-$
	The amount of dissociation of any compound in a solvent is a constant that is determined by the attractions of the ions for each other <i>vs.</i> the attraction between the ions and the solvent.
	In water at 25 °C, the product of the concentrations of H ⁺ and OH ⁻ ions (in $\frac{mol}{l}$) is
	1.0 × 10 ⁻¹⁴ . This number is called the "water dissociation constant" K_w^* . In other words, in water at 25 °C:
	$K_w = [H^+] [OH^-] = 1.0 \times 10^{-14}$
	* K_w is actually the equilibrium constant for the dissociation reaction. $K_{eq} = \frac{[H^+][OH^-]}{[H_2O]}$.
	However, because H_2O is a pure liquid, the concentration of H_2O in pure H_2O is constant— it's just the density divided by the molar mass, which works out to 55.6 M. Therefore, we leave [H_2O] out of the equilibrium expression.

Big Ideas	Details	Unit: Acids & Bases
	Recall that acids create H^+ (or H_3O^+) in water, and b	bases create OH ⁻ in water. In the
	dissociation equation:	
	$H_2O = H^+ + OH$	l [−]
	Le Châtelier's principle predicts that if we add acid, equilibrium to the left, which means $[OH^-]$ decreas $[H^+] [OH^-] = 1.0 \times 10^{-14} = K_w$.	, $[H^+]$ increases. This shifts the ses, and it is still true that
	Similarly, if we add base, $[OH^-]$ increases and $[H^+]$ of $[H^+]$ $[OH^-] = 1.0 \times 10^{-14} = K_w$.	decreases and
	If we have exactly the same amount of acid and baequal to the square root of 1×10^{-14} , which is 1×1 amount of acid and base is said to be <u>neutral</u> .	se, then $[H^+] = [OH^-]$ and both are 0^{-7} M. A solution with the same
	Working with concentrations in scientific notation unwieldy, so we define a function "p" which mean quantity and multiply the result by -1 ." (See the " page 99 for a brief description of the logarithm ma	that vary over 14 powers of ten is s "take the logarithm of the Logarithms" topic starting on thematical function.)
	Therefore, the quantity "pH" would be -log [H ⁺].	
	<u>pH</u> : a measure of the strength of an acidic or basic	solution. Equal to –log [H⁺].
	Examples:	
	if $[H^+] = 0.001$ M, then pH = -log (0.001) = 3 if $[H^+] = 0.000\ 000\ 01$ M (= 1 × 10 ⁻⁸ M) then pH	= -log (1 × 10 ⁻⁸) = 8
	<u>pOH</u> : another measure of the strength of an acidic –log [OH ⁻]. Much less commonly used that pH	c or basic solution. Equal to
	Examples:	
	if $[OH^{-}] = 0.001$ M, then pOH = $-\log (0.001) = 3$ if $[OH^{-}] = 0.000\ 000\ 01$ M (= 1×10^{-8} M) then p	OH = -log (1 × 10 ⁻⁸) = 8
	pH & pOH Equations	
	$pH = -log[H^+]$	pOH = −log[OH ⁻]
	$[H^+] = 10^{-pH}$	[OH ⁻] = 10 ^{-pOH}
	Because the (multiplication) product of [H ⁺][OH ⁻] =	1×10^{-14} , this means that:
	pH + pOH = 14	

Big Ideas	Details				Unit: Acids & Base
	Note that the higher the co	ncentration of H ⁺ ions (I	nigher	value o	of [H ⁺]), the lower
	the pH.				
	Low	v pH = acidic = more H ⁺ :	= less	OH⁻	
	Hig	h pH = basic = less $H^* = I$	more	OH⁻	
	[H⁺]	[OH⁻]	рН	рОН	Acidic/Basic?
	$1 M (= 1 \times 10^{0} M)$	$1 \times 10^{-14} \mathrm{M}$	0	14	very acidic
	$0.1 \text{ M} (= 1 \times 10^{-1} \text{ M})$	$1 \times 10^{-13} \text{M}$	1	13	▲
	0.01 M (= 1 × 10 ⁻² M)	$1 \times 10^{-12} \text{M}$	2	12	
	$1 \times 10^{-3} \mathrm{M}$	$1 \times 10^{-11} \mathrm{M}$	3	11	
	$1 \times 10^{-4} \mathrm{M}$	$1 \times 10^{-10} \mathrm{M}$	4	10	
	$1 \times 10^{-5} \text{M}$	1 × 10 ⁻⁹ M	5	9	slightly acidic
	$1 \times 10^{-6} M$	1 × 10 ⁻⁸ M	6	8	
	$1 \times 10^{-7} \mathrm{M}$	$1 \times 10^{-7} \text{M}$	7	7	neutral
	$1 \times 10^{-8} M$	$1 \times 10^{-6} M$	8	6	
	$1 \times 10^{-9} \mathrm{M}$	1 × 10 ⁻⁵ M	9	5	slightly basic
	$1 \times 10^{-10} \mathrm{M}$	$1 \times 10^{-4} \mathrm{M}$	10	4	▲
	$1 \times 10^{-11} \mathrm{M}$	1 × 10 ⁻³ M	11	3	
	$1 \times 10^{-12} \mathrm{M}$	$0.01 \text{ M} (= 1 \times 10^{-2} \text{ M})$	12	2	
	$1 \times 10^{-13} \mathrm{M}$	$0.1 \text{ M} (= 1 \times 10^{-1} \text{ M})$	13	1	
	$1 \times 10^{-14} M$	$1 M (= 1 \times 10^{0} M)$	14	0	very basic
	Sample Problems:				
	Q: What is the pH of a solu	ition with [H ⁺] = 2.5 x 10) ⁻⁴ M?		
	A: $-\log(2.5 \times 10^{-4}) = 3.60$				
	Q: What is the concentrati	on of H⁺ ions in a solutio	on wit	h a pH o	of 11.4?
	A: $10^{-11.4} = 3.98 \times 10^{-12} \text{ M}$				
	Use this space for summary	and/or additional note	s:		

_

Details Unit: Acids & Bases An aqueous solution is neutral when the concentration of H⁺ and OH⁻ are equal. This occurs in water at pH 7.00 at a temperature of 25 °C. However, remember that temperature affects equilibrium; as the temperature increases, more H⁺ and OH⁻ dissociate. This means [H⁺] and [OH⁻] *both* increase with higher temperatures, which means K_w increases. When that happens, [H⁺] and [OH⁻] are still equal in a neutral solution, but both are larger, and because [H⁺] and [OH⁻] are larger, the pH and pOH are both lower.

Temp. (°C)	Kw	pH of a neutral solution
0	0.114×10^{-14}	7.47
10	0.293×10^{-14}	7.27
20	0.681×10^{-14}	7.08
25	1.008×10^{-14}	7.00
30	1.471×10^{-14}	6.92
40	2.916 × 10 ⁻¹⁴	6.77
50	5.476 × 10 ⁻¹⁴	6.63
100	51.3 × 10 ⁻¹⁴	6.14

In other words, despite what your previous teachers may have taught you, a pH of 7 is only neutral at 25 °C. In fact, in warm-blooded animals with body temperatures around 37 °C, a neutral pH would be approximately 6.8.

This also means that pH + pOH = 14 is only correct at 25 °C.

Use this space for summary and/or additional notes:

Big Ideas

Big Ideas	Details				Unit: A	cids & Bases
	Indicators					
	indicator: a substance that changes color in a specific range of pH values. Indicators					
	are used as a visual way to measure pH.					
	The following table lists some common indicators.					
		Name of Indicator	color in acid	color in base	pH range where color change occurs	
	brom	nophenol blue	e yellow	purple	3.0–4.6	
	meth	nyl red	red	yellow	4.4–6.2	
	litmu	IS	red	blue	5.5-8.2	
	brom	nothymol blue	e yellow	blue	6.0–7.6	
	phen	ol red	yellow	red	6.8–8.4	
	phen	olphthalein	clear	pink	8.2–10.0	
	different color changes over a broader pH range. In fact, some clever chemists have developed a "universal indicator," which is typically composed of water, propanol, phenolphthalein, sodium hydroxide, methyl red, bromothymol blue, and thymol blue. This mixture indicates pH over a range from 3 to 11, in ROYGBIV (rainbow) order:					
		pH range	Descriptior	ו	Color	
		< 3	Strong acid	red		
		3–6	Weak acid	orange	e or yellow	
		7	Neutral	green		
		8–11	Weak base	blue		
		> 11	Strong base	indigo or viol	(dark blue) et (purple)	

Big Ideas	Details Unit: Acids & Bases			
	Homework Problems			
	For each of the following solutions, calculate the information indicated. Choose pH indicators from the "Common Acid-Base Indicators" table in your reference packets.			
	1. $[H+] = 2.5 \times 10^{-4} M$			
	a. pH =			
	b. Is the solution acidic, basic, or neutral?			
	c. Which pH indicator would be best for this solution?			
	2. $[H+] = 4.59 \times 10^{-7} M$			
	a. pH =			
	b. Is the solution acidic, basic, or neutral?			
	c. Which pH indicator would be best for this solution?			
	3. pH = 9.1			
	a. [H+] =			
	b. Is the solution acidic, basic, or neutral?			
	c. Which pH indicator would be best for this solution?			

Big Ideas Deta	ils	Unit: Acids & E
2	. pH = 5	.5
	a.	[H+] =
	b.	Is the solution acidic, basic, or neutral?
	C.	Which pH indicator would be best for this solution?
5	6. [OH–]	= 7.9 × 10 ⁻⁷ M
	a.	[H+] =
	b.	рН =
	C.	Is the solution acidic, basic, or neutral?
	d.	Which pH indicator would be best for this solution?

	pKa & Buffers	Page: 504
as	Details	Unit: Acids & Bases
	pK _a & Buffers	
	Unit: Acids & Bases	
	MA Curriculum Frameworks (2016): N/A	
	Mastery Objective(s): (Students will be able to)	
	• Calculate pH from $[H^+]$ and pOH from $[OH^-]$.	
	 Identify acids and bases from their pK_a values. 	
	 Select an appropriate indicator for a desired pH ran 	ige.
	Success Criteria:	•
	 pH and pOH are calculated correctly. 	
	 Acids and bases are correctly identified from their p 	oKa values.
	• Indicator changes color in a pH range that includes	the pH of the given acid or
	base.	
	Tier 2 Vocabulary: acid, base, indicator	
	Language Objectives:	
	 Explain why higher [H⁺] results in a lower pH. 	
	Notes:	
	Acid-base chemistry is largely equilibrium chemistry in wh	nich the solvent, usually
	H ₂ O, plays a significant role.	, ,
	As stated earlier, water dissociates into H ⁺ and OH ⁻ ions.	Acids and bases change
	the concentrations of H ⁺ and OH ⁻ ions in solution, which c	can have significant effects
	acid dissociation constant (K _a): is the equilibrium constant acid. For the "generic" acid HA:	t for the dissociation of an
	[H ⁺][A ⁻]	
	$\kappa_a =$	
	The greater the K_a value, the stronger the acid. (Rem	ember your negative
	exponents! <i>E.g.</i> , 10^{-5} is <u>greater</u> than 10^{-7} .)	. –
	$\underline{pK_a} = -\log K_a$ (analogous to pH). The lower (or more neg	ative) the p K_a , the
	stronger the acid.	
	When exactly 50 % of the acid HA is neutralized, [HA]	= [A⁻], and the above
	formula reduces to $K_a = [H^+]$. This means that pH = p <i>I</i>	K _a when the acid is half-
	neutralized.	
pKa & Buffers

Big Ideas	Details Unit: Acids & Bases
	base dissociation constant (K _b): is the equilibrium constant for the dissociation of a base. For the "generic" base B:
	$K_{b} = \frac{[HB^{+}][OH^{-}]}{[B]}$
	We can use the concept of pK_a to add to our definitions of strong acids and bases:
	strong acid: an acid with a pK_a lower than that of H_3O^+ (1.0). Strong acids include HCl, HBr, HI, H_2SO_4 and HNO ₃ .
	Strong acids dissociate completely into H^+ and the corresponding anion. The dissociated H^+ converts H_2O molecules to H_3O^+ ions.
	weak acid: an acid with a p K_a higher than that of H ₃ O ⁺ (1.0), but less than 7.0 (the pH of a neutral solution at 25 °C).
	<u>strong base</u> : a base whose conjugate acid is weaker than H_2O (<i>i.e.</i> , whose conjugate acid has a pK _a higher than 14). Hydroxides are strong bases because they release OH ⁻ . However, note that aqueous Mg(OH) ₂ <u>acts</u> more like a weak base because the limited solubility of Mg(OH) ₂ results in a concentration of OH ⁻ that is similar to that produced by a weak base.
	Strong bases either release OH ⁻ ions directly into solution, or form OH ⁻ ions by pulling H ⁺ off of H ₂ O molecules.
	weak base: a base whose conjugate acid has a pK_a higher than 7.0 but less than 14.

Use this space for summary and/or additional notes:

pKa & Buffers

Big Ideas	Details Unit: Acids & Bases
	Buffers
	<u>buffer</u> : a weak acid or base that prevents the pH of a solution from changing drastically until it neutralizes the buffer.
	For example, if you have a fish tank, you want to keep the pH from getting too low, you could add NaHCO ₃ . The reaction:
	$H^+ + HCO_3^- \longrightarrow H_2CO_3$
	occurs around pH 6.4. As acid accumulates in your fish tank, it will react with the HCO_3^- ions, and the pH will remain above 6.4 until all of the HCO_3^- ions have been converted to H_2CO_3 .
	Buffers can work in either direction—to absorb acid or base. If you use a combination of two buffers (one above and one below your desired pH), you can keep the pH within a narrow range.
	In fact, water acts as a buffer, but over a very wide pH range. The pH of an aqueous solution is limited, because stronger acids just convert more H_2O to H_3O^+ , and stronger bases just convert more H_2O to OH^- . The presence of water effectively keeps the pH between 1 and 14. In fact, the reason your biology teacher taught you that the pH range goes from 1–14 is because acid-base reactions in biology all happen in aqueous environments.

Use this space for summary and/or additional notes:

		pKa & Buffers	Page: 507			
Big Ideas	Details		Unit: Acids & Bases			
		Homework Problems				
	 Rank the following acids from strongest to weakest, based on their pKa values. Refer to of your Chemistry Reference Tables. 					
	Based o your Ch the stro	Based on pKa values in "Table P. <i>pKa</i> Values for Common Acids" on page 513 of your Chemistry Reference Tables, rank the following ten compounds in order, from the strongest base.				
		HF, HCN, HCl, HPO_4^{2-} , HNO_3 , H_2O , CH_3COOH , NH_4^+ ,	H ₂ SO ₄ , H ₂ CO ₃			
	1.	5	8			
	2.	6	9			
	3.	7	10			
	4.					
	11.	The wastes from fish in a fish tank produce acids, whi water in the tank to decrease over time. Which acid- table of <i>pKa</i> Values for Common Acids would be most the pH from dropping below 7.0. Explain.	ch cause the pH of the base pair from the effective at keeping			
	Use this	s space for summary and/or additional notes:				

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Table A. Standard Temperature and Pressure				
Name Values				
"Standa	rd" Pressure	1 atm	760 torr	101 . 3 kPa
Standar	d Temperature	0 °C	32°F	273.15K
atm =	atmosphere			
Torr =	millimeter of mercury (mm Hg)			
kPa =	kilopascal			
°C =	degree Celsius			
°F =	degree Fahrenh	neit		
К =	kelvin			

Table C. Selected Prefixes					
Factor	Number of Units	Prefix	Symbol		
10 ⁶	1,000,000	mega-	Μ		
10 ³	1,000	kilo-	k		
10 ⁻¹	0.1	deci-	d		
10 ⁻²	0.01	centi-	с		
10 ⁻³	0.001	milli-	m		
10 ⁻⁶	0.000 001	micro-	🛛 (or u)		
	1				

Table B. Selected Units		
Name	Symbol	Quantity
meter (SI)	m	longth
centimeter	cm	length
kilogram (SI)	kg	mass
gram	g	111855
Pascal (SI derived)	Ра	
atmosphere	atm	proceuro
mm of mercury	mm Hg	pressure
Torr	Torr	
Kelvin (SI)	К	tomporaturo
degree Celsius	°C	temperature
amt of substance (SI)	mol	mole
Joule (SI derived)	J	0.000
kilocalorie	kcal	energy
second (SI)	S	time
liter	L, €	volume
part per million	ppm	concentration
molarity	M, $\frac{\text{mol}}{\ell}$	concentration

Table D. Physical Constants for Water	
Freezing Point @ 1 atm	0 °C = 273.15 K
Boiling Point @ 1 atm	100 °C = 373.15 K
Heat of Fusion	333 . 6 J/g
Heat of Vaporization	2270 J/g
Specific Heat Capacity (C_{ρ})	4 . 184 J/g·°C
Freezing Point Depression Constant (K _f)	0.52 °C/m
Boiling Point Elevation Constant (K _b)	1.86 °C/ <i>m</i>

Table E. Vapor Pressure and			
Temp (°C)	P _{vap} (kPa)	density (g/cm³)	
0.01	0.61173	0.99978	
1	0.65716	0.99985	
4	0.81359	0.99995	
5	0.87260	0.99994	
10	1.2281	0.99969	
15	1.7056	0.99909	
20	2.3388	0.99819	
25	3.1691	0.99702	
30	4.2455	0.99561	
35	5.6267	0.99399	
40	7.3814	0.99217	
45	9.5898	0.99017	
50	12.344	0.98799	
55	15.752	0.98565	
60	19.932	0.98316	
65	25.022	0.98053	
70	31.176	0.97775	
75	38.563	0.97484	
80	47.373	0.97179	
85	57.815	0.96991	
90	70.117	0.96533	
95	84.529	0.96192	
100	101.32	0.95475	
105	120.79	0.95475	



Chemistry 1

Table G. Solubility Guidelines					
Ions That Form		Ions That Form INSOLUBLE			
SOLUBLE Compounds		Compounds			
Group I ions (Li ⁺ , Na ⁺ , etc.)		carbonate (CO ₃ ^{2–})			
ammonium (NH4 ⁺)		chromate (CrO ₄ ^{2–})	Group I ions,		
nitrate (NO₃ [−])		phosphate (PO4 ^{3–})	ammonium (NH ₄ ⁺)		
hydrogen carbonate (HCO₃⁻)		sulfite (SO₃²-)			
chlorate (ClO₃ [–])		$cultide (S^2)$	Group I ions,		
perchlorate (ClO ₄ ⁻)		sunde (S)	Group II ions, NH4 ⁺		
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	Ag⁺	hydroxide (OH⁻)	Group I ions,		
halides (Cl⁻, Br⁻, I⁻)	Ag ⁺ , Cu ⁺ , Pb ²⁺ , Hg ₂ ²⁺	oxide (O ^{2–})	NH4 ⁺ , Ba ²⁺ , Sr ²⁺ , Tl ⁺		
sulfates (SO4 ^{2–})	Ca ²⁺ , Sr ²⁺ , Ba ²⁺ , Ag ⁺ , Pb ²⁺				

Table H. K _{sp} Values for Some				
insoluble Salts at 25 °C				
Compound	Ksp			
MgCO ₃	1.0×10^{-5}			
PbCl ₂	1.7 × 10 ⁻⁵			
BaF ₂	2.0 × 10 ⁻⁶			
CuCl	1.0×10^{-6}			
Pbl ₂	1.6×10^{-8}			
AgOH	1.0×10^{-8}			
BaCO₃	8.1 × 10 ⁻⁹			
CaCO₃	3.8 × 10 ⁻⁹			
SrCO₃	9.4 × 10 ⁻¹⁰			
AgCl	1.8×10^{-10}			
BaSO ₄	1.1×10^{-10}			
CaF ₂	3.9×10^{-11}			
Mg(OH)₂	1.0×10^{-11}			
Ag ₂ CrO ₄	9.0×10^{-12}			
Cul	5.0×10^{-12}			
AgBr	3.3×10^{-13}			
PbSO ₄	2.5×10^{-13}			
PbCO₃	1.6×10^{-13}			
Mn(OH)₂	4.0×10^{-14}			
PbCrO ₄	1.8×10^{-14}			
Fe(OH)₂	1.6×10^{-14}			
Agl	1.5×10^{-16}			
Zn(OH)₂	7.9×10^{-18}			
FeS	4.0×10^{-18}			
HgCl	2.0×10^{-18}			
ZnS	1.0×10^{-23}			
PbS	8.4 × 10 ⁻²⁸			
CdS	3.6 × 10 ⁻²⁹			
Al(OH)₃	1.6×10^{-34}			
CuS	8.7 × 10 ⁻³⁶			
Fe(OH)₃	1.3×10^{-36}			
Ag ₂ S	2.0×10^{-50}			
HgS	3.0×10^{-53}			



Table J. Number Prefixes						
Number	Inorganic	Organic	Number	Inorganic	Organic	
1	mono-	meth-	6	hexa-	hex-	
2	di-	eth-	7	hepta-	hept-	
3	tri-	prop-	8	octa-	oct-	
4	tetra-	but-	9	nona-	non-	
5	penta-	pent-	10	deca-	dec-	

Table K. Polyatomic lons							
ion	formula	ion	formula	ion	formula	ion	formula
americyl	AmO ₂ ²⁺	ammonium	NH_4^+	cyanate	OCN [−]	dichromate	Cr ₂ O ₇ ²⁻
carbonyl	CO ²⁺	hydronium	H₃O⁺	thiocyanate	SCN⁻	imide	NH ²⁻
thiocarbonyl	CS ²⁺	iodyl	10_2^+	selenocyanate	SeCN⁻	molybdate	MoO4 ²⁻
chromyl	CrO ₂ ²⁺	nitrosyl	NO ⁺	tellurocyanate	TeCN⁻	peroxide	O ₂ ²⁻
neptunyl	NpO ₂ ²⁺	thionitrosyl	NS⁺	hydroxide	OH⁻	oxalate	$C_2O_4^{2-}$
plutoryl	PuO ₂ ²⁺	phosphoryl	PO⁺	iodate	IO₃ [−]	phthalate	$C_8H_4O_4^{2-}$
selinyl	SeO ²⁺	thiophosphoryl	PS⁺	methanolate	CH₃O⁻	selenate	SeO4 ²⁻
selenoyl	SeO ₂ ²⁺	phospho	PO ₂ ⁺	methanethiolate	CH₃S⁻	disulfide	S_2^{2-}
thionyl / sulfinyl	SO ²⁺	acetate	CH₃COO⁻	ethanolate	C₂H₅O [−]	sulfate	SO4 ²⁻
sulfonyl / sulfuryl	SO2 ²⁺	amide	NH2 ⁻	permanganate	MnO₄⁻	thiosulfate	$S_2O_3^{2-}$
uranyl	UO ²⁺	hydroxylamide	NHOH ⁻	nitrate	NO₃⁻	dithionate	$S_2O_4^{2-}$
vanadyl	VO ²⁺	azide	N ₃ ⁻	superoxide	O ₂ ⁻	silicate	SiO ₃ ²⁻
mercury (II)	Hg ²⁺	hydrazide	$N_2H_3^-$	tetraborate	B4O7 ²⁻	borate	BO₃ ^{3−}
mercury (I)	Hg_{2}^{2+}	bromate	BrO₃⁻	carbide	C_2^{2-}	arsenate	AsO ₄ ³⁻
		chlorate	ClO3 [−]	carbonate	CO2 ²⁻	phosphate	PO4 ³⁻
		cyanide	CN⁻	chromate	CrO ₄ ²⁻	orthosilicate	SiO4 ⁴⁻

Table L. Flame Test Colors					
Element	Color	Element	Color	Element	Color
Ba	yellow-green	К	pink	Pb	blue
Са	orange-red	Li	fuchsia	Sb	pale green
Cu	blue-green	Mg	bright white	Sr	red
Fe	gold	Na	yellow	Zn	blue-green

Table M. Aqueous Ion Colors				
lon	Color	lon	Color	
Cu⁺	green	V ²⁺	violet	
Cu ²⁺	blue	V ³⁺	blue-green	
Fe ²⁺	yellow-green	CrO ₄ ^{2–}	yellow	
Fe ³⁺	orange-red	Cr ₂ O ₇ ²⁻	orange	
Cr ³⁺	violet [Cr(NO₃)₃] to green [CrCl₃]	Cu(NH ₃) ₄ ²⁺	dark blue	
Ni ²⁺	green	FeSCN ²⁺	red-brown (wine-red to dark orange)	
Mn ²⁺	pink	Co ²⁺	pink	
Mn ⁷⁺	purple (<i>e.g.,</i> the MnO ₄ ⁻ ion)	CoCl ₄ ^{2–}	blue	
Pb ³⁺	blue-green (Pb ²⁺ and Pb ⁴⁺ are clear)	Ti(H ₂ O) ₆ ³⁺	purple	

Table N. Colors of Assorted Compounds				
Compound	Color	Compound	Color	
F ₂	pale yellow gas	NO	colorless gas	
Cl ₂	green-yellow gas	NO ₂	brown gas	
Br ₂	red-brown liquid	metallic	sulfides of transition metals	
l ₂	dark metallic solid; dark violet vapor	sulfides	tend to be black	
S ₈	yellow odorous solid	metallic	oxides of colored transition metals	
Pbl ₂	bright yellow precipitate	oxides	tend to be colored	
Fe ₂ O ₃	reddish-brown (rust)			

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Table O. Common Acids			
Formula	Name		
HCl (aq)	hydrochloric acid		
HNO₃ (aq)	nitric acid		
H ₂ SO ₄ (aq)	sulfuric acid		
H ₃ PO ₄ (<i>aq</i>)	phosphoric acid		
H ₂ CO ₃ (aq)	carbonic acid		
HC₂H₃O₂ <i>(aq)</i> or CH₃COOH <i>(aq)</i>	ethanoic acid (acetic acid)		

Table Q. Common Bases	
Formula	Name
NaOH (aq)	sodium hydroxide
КОН (аq)	potassium hydroxide
Ca(OH) ₂ <i>(aq)</i>	calcium hydroxide
NH₃ <i>(aq)</i>	aqueous ammonia

Table R. Common Acid-Base Indicators				
Indicator	pH Range of Color Change	Color Change		
bromophenol blue	3.0-4.6	yellow-purple		
methyl orange	3.2 – 4.4	red-yellow		
bromocresol green	3.8 – 5.4	yellow-blue		
methyl red	4.4 - 6.2	red-yellow		
litmus	5.5 – 8.2	red-blue		
bromothymol blue	6.0 - 7.6	yellow-blue		
phenol red	6.8 - 8.4	yellow-red		
thymol blue	8.0 - 9.6	yellow-blue		
phenolphthalein	8.2 – 10	clear-pink		

Table P. <i>p</i> K _a Values for Common Acids				
Acid	pKa	Conj. Base		
H ₂ O	15.7	OH⁻		
HPO4 ²⁻	12.6	PO4 ³⁻		
HCO₃⁻	10.2	CO3 ²⁻		
NH_4^+	9.2	NH₃		
HCN	9.1	CN⁻		
$H_2PO_4^-$	7.2	HPO4 ²⁻		
H ₂ S	7.0	HS⁻		
H ₂ CO ₃	6.4	HCO₃⁻		
CH₃COOH	4.8	CH₃COO⁻		
нсоон	3.7	HCOO ⁻		
HNO ₂	3.3	NO ₂ ⁻		
HF	3.2	F⁻		
C ₆ H ₈ O ₇ (citric acid)	3.1	$C_6H_7O_7^-$		
H ₃ PO ₄	2.2	H₂PO₄ [−]		
HSO₄ [−]	2.0	SO4 ²⁻		
HNO₃	-1.4	NO₃ [−]		
H₃O ⁺	-1.7	H ₂ O		
HCI	-7.0	Cl⁻		
HBr	-9.0	Br⁻		
HI	-10	I-		
HCIO ₄	-10	CIO4 ⁻		
H_2SO_4	-12	HSO₄ [−]		

Any acid with a pK_{α} value less than 1 is a strong acid; any base whose conjugate acid has a pK_{α} value greater than 14 is a strong base.



Table T. Symbols Used in Nuclear Chemistry			
Name	Notation	Symbol	
alpha particle	4_2 He or ${}^4_2\alpha$	α	
beta particle (electron)	$^{0}_{-1}e$ or $^{0}_{-1}eta$	β^-	
gamma radiation	${}^0_0 \gamma$	γ	
neutron	${}^{1}_{0}n$	n	
proton	$^1_1\mathrm{H} { m or} ^1_1 p$	p	
positron	$^{0}_{+1}e$ or $^{0}_{+1}eta$	β^+	

Table V. Constants Used in Nuclear Chemistry			
Constant Value			
mass of an electron (m _e)	0.00055 amu		
mass of a proton (m_p)	1.00728 amu		
mass of a neutron (m_n)	1.00867 amu		
Becquerel (Bq)	1 disintegration/second		
Curie (Ci)	3.7 x 10 ¹⁰ Bq		



Table U. Selected Radioisotopes					
Nuclide	Half-Life	Decay Mode			
³ Н	12.26 y	β-			
¹⁴ C	5730 y	β-			
¹⁶ N	7.2 s	β-			
¹⁹ Ne	17.2 s	β^+			
²⁴ Na	15 h	β^-			
²⁷ Mg	9.5 min	β^-			
³² P	14.3 d	β-			
³⁶ Cl	3.01 × 10 ⁵ y	β-			
³⁷ K	1.23 s	β^+			
⁴⁰ K	1.26 × 10 ⁹ у	β^+			
⁴² K	12.4 h	β^-			
³⁷ Ca	0.175 s	β^-			
⁵¹ Cr	27.7 d	α			
⁵³ Fe	8.51 min	β-			
⁵⁹ Fe	46.3 d	β^-			
⁶⁰ Co	5.26 y	β^-			
⁸⁵ Kr	10.76 y	β^-			
⁸⁷ Rb	$4.8 \times 10^{10} \text{ y}$	β-			
⁹⁰ Sr	28.1 y	β-			
⁹⁹ Tc	2.13 x 10⁵ y	β^-			
¹³¹	8.07 d	β^-			
¹³⁷ Cs	30.23 y	β^-			
¹⁵³ Sm	1.93 d	β^-			
¹⁹⁸ Au	2.69 d	β^-			
²²² Rn	3.82 d	α			
²²⁰ Fr	27.5 s	α			
²²⁶ Ra	1600 y	α			
²³² Th	1.4 x 10 ¹⁰ y	α			
²³³ U	1.62 x 10⁵ y	α			
²³⁵ U	7.1 x 10 ⁸ y	α			
²³⁸ U	4.51 x 10 ⁹ y	α			
²³⁹ Pu	2.44 x 10 ⁴ y	α			
²⁴¹ Am	432 y	α			

Table X. Activity Series	Table X. Activity Series					
Metals		Non-				
	metals					
	Cs	▲ F ₂				
	Rb	Cl ₂				
Reacts with cold H ₂ O	к	Br ₂				
and acids, replacing	Na	· I2				
hydrogen. Reacts with	Li					
O ₂ , forming oxides.	Ва					
	Sr					
	Ca					
	Mg					
Departs with steam (not	Be					
Reacts with steam (not	Al					
	Mn					
replacing hydrogen.	Zn					
Reacts with O_2 ,	Cr					
forming oxides.	Fe					
	Cd					
Does not react with H ₂ O.	Со					
Reacts with acids,	Ni					
replacing hydrogen.	Sn					
Reacts with O ₂ ,	Pb					
forming oxides.	H ₂					
Boasts with Os forming	Sb					
avidas	Bi					
oxides.	Cu					
Eairly uproactive	Ag					
forming ovides only	Hg					
indirectly	Au					
indirectly.	Pt					

Arrows indicate direction from LEAST to MOST active elements. An element can replace any element below itself on the activity series.

Note that the order of elements in the activity series is similar to, though not quite identical with, the order of elements/ions in the table of Std. Reduction Potentials (Table Y).

Table Y. Std. Reduction Potentials						
Half-Reaction	E ⁰ (V)					
$Li^+ + e^- \leftrightarrows Li(s)$	-3.040					
$Cs^+ + e^- \leftrightarrows Cs(s)$	-3.026					
$Rb^+ + e^- \leftrightarrows Rb(s)$	-2.98					
$K^+ + e^- \leftrightarrows K(s)$	-2.931					
$Ba^{2+} + 2e^{-} \leftrightarrows Ba(s)$	-2.912					
$Sr^{2+} + 2e^{-} \leftrightarrows Sr(s)$	-2.899					
$Ca^{2+} + 2e^{-} \leftrightarrows Ca(s)$	-2.868					
Na⁺ + e⁻ 与 Na(s)	-2.71					
$Mg^{2+} + 2e^{-} \leftrightarrows Mg(s)$	-2.372					
$Be^{2+} + 2e^{-} \leftrightarrows Be(s)$	-1.85					
$AI^{3+} + 3e^{-} \leftrightarrows AI(s)$	-1.66					
Mn ²⁺ + 2 <i>e</i> [−] ≒ Mn(<i>s</i>)	-1.029					
2 H₂O + 2 <i>e</i> [−] ≒	-0.828					
$H_2(g) + 2 OH^-$	0.020					
$Zn^{2+} + 2e^{-} \leftrightarrows Zn(s)$	-0.762					
$Cr^{3+} + 3e^{-} \leftrightarrows Cr(s)$	-0.74					
Fe ²⁺ + 2 <i>e</i> ⁻ ≒ Fe(<i>s</i>)	-0.44					
$Cr^{3+} + e^{-} \leftrightarrows Cr^{2+}$	-0.42					
$Cd^{2+} + 2e^{-} \leftrightarrows Cd(s)$	-0.40					
$\operatorname{Co}^{2+} + 2e^{-} \leftrightarrows \operatorname{Co}(s)$	-0.28					
$Ni^{2+} + 2e^{-} \leftrightarrows Ni(s)$	-0.25					
$\operatorname{Sn}^{2+} + 2e^{-} \leftrightarrows \operatorname{Sn}(s)$	-0.13					
$Pb^{2+} + 2e^{-} \leftrightarrows Pb(s)$	-0.13					
$2 H^+ + 2 e^- \leftrightarrows H_2(g)$	0.000					
$S(s) + 2 H^+ + 2 e^- \leftrightarrows H_2 S(g)$	+0.14					
$\operatorname{Sn}^{4+} + 2e^{-} \leftrightarrows \operatorname{Sn}^{2+}$	+0.15					
$Cu^{2+} + e^{-} \leftrightarrows Cu^{+}$	+0.159					
$\operatorname{Cu}^{2+} + 2e^{-} \leftrightarrows \operatorname{Cu}(s)$	+0.340					
$Cu^+ + e^- \leftrightarrows Cu(s)$	+0.520					
$I_2(s) + 2e^- \leftrightarrows 2I^-$	+0.54					
$Fe^{3+} + e^{-} \leftrightarrows Fe^{2+}$	+0.77					
$Ag^+ + e^- \leftrightarrows Ag(s)$	+0.800					
$Hg_2^{2+} + 2e^{-} \leftrightarrows 2Hg(I)$	+0.80					
$Hg^{2+} + 2e^{-} \leftrightarrows Hg(I)$	+0.85					
$2 \operatorname{Hg}^{2+} + 2 e^{-} \leftrightarrows \operatorname{Hg}_{2^{2+}}$	+0.91					
Br₂(/) + 2 e ⁻ ≒ 2 Br ⁻	+1.07					
$Pt^{2+} + 2e^{-} \leftrightarrows Pt(s)$	+1.188					
$O_2(g) + 4 H^+ + 4 e^- \leftrightarrows 2 H_2O$	+1.23					
$Cl_2(g) + 2e^- \leftrightarrows 2Cl^-$	+1.36					
Au ³⁺ + 3 <i>e</i> ⁻ 与 Au(<i>s</i>)	+1.52					
$Co^{3+} + e^{-} \leftrightarrows Co^{2+}$	+1.82					
F₂(g) + 2 e ⁻ ≒ 2 F ⁻	+2.87					

F

Table Z. Selected Properties of the Elements										
			atomic	melting	hoiling	specific		electro-	1 st ioniza-	common
atomic	atomic	element	mass	noint	noint	heat	density	neg-	tion	oxidation
#	symbol	name	(IUPAC	°C	°C	capacity	(g/mL)	ativity	potential	states
			2005)		Ŭ	(J/g∙K)		(Pauling)	(kJ/mol)	States
89	Ac	actinium	227	1050	3200	—	10.1	1.1	499	+3
13	Al	aluminum	26.98	660	2467	0.9	2.7	1.61	578	+3
95	Am	americium	243	994	2607	0.11	13.7	1.3	578	+3,4,5,6
51	Sb	antimony	121.8	631	1950	0.21	6.69	2.05	834	+3,5
18	Ar	argon	39.95	-189.2	-185.7	0.52	0.00178		1521	0
33	As	arsenic	74.92	817	617	0.33	5.73	2.18	947	±3,+5
85	At	astatine	210	302	337	—	_	2.2	917	
56	Ва	barium	137.3	725	1640	0.204	3.5	0.89	503	+2
97	Bk	berkelium	247	986	—	—	14	1.3	601	+3,4
4	Be	beryllium	9.012	1278	2970	1.82	1.85	1.57	899	+2
83	Bi	bismuth	209.0	271	1560	0.12	9.75	2.02	703	+3,5
107	Bh	bohrium	272	_	_	_	_	_	_	_
5	В	boron	10.81	2079	2550	1.02	2.34	2.04	801	+3
35	Br	bromine	79.90	-7.2	58.8	0.473	3.12	2.96	1140	±1,+5
48	Cd	cadmium	112.4	320.9	765	0.23	8.65	1.69	868	+2
20	Са	calcium	40.08	839	1484	0.63	1.55	1	590	+2
98	Cf	californium	251	_	_	_	_	1.3	608	+3
6	С	carbon	12.01	3367	4827	0.71	2.25	2.55	1086	±4
58	Ce	cerium	140.1	798	3257	0.19	6.66	1.12	534	+3.4
55	Cs	cesium	132.9	28.4	669	0.24	1.87	0.79	376	+1
17	Cl	chlorine	35.45	-101	-34.6	0.48	0.00321	3.16	1251	-1
24	Cr	chromium	52.00	1857	2672	0.45	7.19	1.66	653	+3.2.6
27	Co	cobalt	58.93	1495	2870	0.42	8.9	1.88	760	+2.3
112	Cn	copernicum	285	_		_	_	_	_	
29	Cu	conner	63 55	1083	2567	0 38	8 96	19	745	+2 1
96	Cm	curium	247	1340			13 5	13	581	+3
110	Ds	darmstadtium	247	-	_	_			_	
105	Dh	dubnium	268		_	_	_	_	_	_
66	Dv	dysprosium	162.5	1/17	2567	0 17	8 5 5	1 22	573	+3
99	Fs	einsteinium	252		2507		0.55	1 3	619	+3
68	Er Er	orbium	167.3	1520	2868	0.17	9.07	1.5	589	+3
63	Eu	europium	152.0	822	1520	0.17	5.07	1.24	547	+3.2
100	Em	formium	257	- 022	1525	0.10	J.24	1 2	627	+3,2
114		florovium	20					1.5	027	-
0	с ГІ	fluorino	10.00	_210.9	_100 1	0.82	0.0017	2 00	1691	_1
9	Г Бr	francium	19.00	-219.0	-100.1	0.82	0.0017	5.90	200	-1
6/			157.2	1212	2272	0.22	7.0	0.7	500	+1
64 21	Ga	gadolinium	157.3	1313	32/3	0.23	7.9	1.2	593	+3
31	Ga	gailium	69.72 72.62	29.8	2403	0.37	5.9	1.81	579	+3
32	Ge	germanium	/2.63	947.4	2830	0.32	5.32	2.01	762	+4,2
/9	AU	gold	197.0	1064	3080	0.128	19.3	2.54	890	+3,1
12	HT	nathium	1/8.5	2227	4600	0.14	13.3	1.3	659	+4
108	<u>Hs</u>	nassium	270			-		_		
2	не	nellum	4.003	-2/2.2	-268.9	5.193	0.000179	4 22	2372	0
6/	Но	holmium	164.9	14/4	2700	0.16	8.8	1.23	581	+3
1	<u> </u>	hydrogen	1.008	-259.1	-252.9	14.304	0.0000699	2.2	1312	±1
49	In	indium	114.8	156.6	2080	0.23	7.31	1.78	558	+3
53	l ,	iodine	126.9	113.5	184	0.214	4.93	2.66	1008	-1,+5,7
77	lr	iridíum	192.2	2410	4130	0.13	22.4	2.2	878	+4,3,6
26	Fe	iron	55.85	1535	2750	0.44	7.86	1.83	762	+3,2
36	Kr	krypton	83.80	-157	-152	0.248	0.00374	3	1351	0
57	La	lanthanum	138.9	920	3454	0.19	6.15	1.1	538	+3
103	Lr	lawrencium	262	_	_	_	—	_	_	+3
82	Pb	lead	207.2	327.5	1740	0.13	11.4	2.33	716	+2,4
3	Li	lithium	6.968	180.5	1342	3.6	0.543	0.98	520	+1
116	Lv	livermorium	293	—	—	—	—	—	—	—
71	Lu	lutetium	175.0	1663	3402	0.15	9.84	1.27	524	+3
12	Mg	magnesium	24.31	649	1090	1.02	1.74	1.31	738	+2
25	Mn	manganese	54.94	1244	1962	0.48	7.43	1.55	717	+2,3,4,6,7

			atomic			specific		electro-	1 st ioniza-	
atomic	atomic	element	mass	melting	boiling	heat	density	neg-	tion	common
#	symbol	name		point	point	canacity	(g/ml)	ativity	notential	oxidation
п	Symbol	name	200E)	°C	°C		(5/1112)	(Douling)	(kl/mol)	states
100			2005)			(J/g.k)		(Pauling)	(KJ/MOI)	
109	Mt	meitnerium	276	—	—	_	_	_	_	_
101	Md	mendelevium	258	—	—	—	_	1.3	635	+3,2
80	Hg	mercury	200.6	-38.9	357	0.139	13.5	2	1007	+2,1
42	Мо	molybdenum	95.95	2617	4612	0.25	10.2	2.16	684	+6,3,5
115	Mc	moscovium	288	_	_	_	_	_	_	_
60	Nd	neodvmium	144.2	1016	3127	0.19	7	1.14	533	+3
10	Ne	neon	20.18	-2/18	-2/18 7	0.904	0.0009		2081	0
02	No	nontunium	20.10	640	2000	0.504	20.2	1 26	605	+5246
20	NP NI:	neptunium	237	1452	2720	0.12	20.2	1.50	205	+3,3,4,0
28	INI	піскеі	58.69	1453	2730	0.44	8.9	1.91	/3/	+2,3
113	Nh	nihonium	284	_	_	_	_	_	_	—
41	Nb	niobium	92.91	2468	4742	0.26	8.57	1.6	652	+5,3
7	Ν	nitrogen	14.01	-209.9	-195.8	1.04	0.00125	3.04	1402	-3
102	No	nobelium	259	—	_	—	_	1.3	642	+2,3
118	Og	oganesson	294	_	_	_	_	_	_	_
76	Os	osmium	190.2	3045	5030	0.13	22.6	2.2	839	+4.6.8
2 2	0	oxygen	16.00	_218 /	_183	0.92	0.001/13	3 1/	131/	-2
16	Dd	nalladium	10.00	10.4	2140	0.92	12	3.44	204	12.4
40	Pu	panaulum	20.07	1554	200	0.24	1.02	2.2	004 1012	+2,4
15	P	pnospnorus	30.97	44.1	280	0.77	1.82	2.19	1012	-3
78	Pt	platinum	195.1	1772	3827	0.13	21.4	2.28	868	+4,2
94	Pu	plutonium	244	641	3232	0.13	19.8	1.28	585	+4,3,5,6
84	Ро	polonium	209	254	962	0.12	9.32	2	812	+4,2
19	К	potassium	39.10	63.25	760	0.75	0.86	0.82	419	+1
59	Pr	praseodymium	140.9	931	3017	0.19	6.77	1.13	527	+3,4
61	Pm	promethium	145	1042	3000	0.18	7.26	_	535	+3
01	 Do	protactinium	221.0	1570	4000	0.12	15.4	15	568	+5 /
00	Pa	radium	231.0	700	1140	0.12	1J.4 E	1.5	508	· J,+
00	Г.d D.:	raulum	220	700	1140	0.12	5	0.9	509	+2
80	RN	radon	222	-/1	-61.8	0.09	0.00973		1037	0
75	Re	rhenium	186.2	3180	5600	0.13	21	1.9	760	+7,4,6
45	Rh	rhodium	102.9	1966	3727	0.242	12.4	2.28	720	+3,4,6
111	Rg	roentgentium	280	—	—	_	—	_	_	_
37	Rb	rubidium	85.47	38.9	686	0.363	1.53	0.82	403	+1
44	Ru	ruthenium	101.1	2310	3900	0.238	12.4	2.2	710	+4,3,6,8
104	Rf	rutherfordium	267	_	_	_	_	_	_	
62	Sm	samarium	150 /	107/	170/	0.2	7 5 2	1 17	545	T 3 0
21	Sill	sandium	110.4	1074	1/34	0.2	2.00	1.17	622	13,2
100	SC C-	Scanului	44.90	1541	2052	0.0	2.99	1.50	055	+5
100	Sg	seaborgium	2/1				_			_
34	Se	selenium	78.97	217	685	0.32	4.79	2.55	941	+4,-2,+6
14	Si	silicon	28.09	1410	2355	0.71	2.33	1.9	787	±4
47	Ag	silver	107.9	962	2212	0.235	10.5	1.93	731	+1
11	Na	sodium	22.99	97.8	883	1.23	0.971	0.93	496	+1
38	Sr	strontium	87.62	769	1384	0.3	2.54	0.95	549	+2
16	S	sulfur	32.07	112.8	444.7	0.71	2.07	2.58	1000	-2
73	Та	tantalum	180.9	2996	5425	0 14	16.6	15	761	+5
43	Tc	technetium	98	2172	4877	0.21	11 5	1 9	702	+746
50	To	tollurium	127 6	710 5	000 0	0.21	6.24	1.J 2 1	860	·,,-,0 ±16_0
32	ास म	tenulium	127.0	443.3	202.0	0.2	0.24	۲.٦	605	±4,0,=∠
11/	1 S 	tennessine	292	-	_	_	_	—	_	_
65	ID	terbium	158.9	1365	3230	0.18	8.23	_	569	+3,4
81	TI	thallium	204.4	303	1457	0.13	11.9	1.62	589	+1,3
90	Th	thorium	232.0	1750	4790	0.12	11.7	1.3	587	+4
69	Tm	thulium	168.9	1545	1950	0.16	9.32	1.25	597	+3,2
50	Sn	tin	118.7	232	2270	0.227	7.31	1.96	709	+4,2
22	Ti	titanium	47.87	1660	3287	0.52	4.54	1.54	659	+4.32
7/	\\/	tungsten	182.8	2/10	5660	0.12	10 2	2 26	770	+6 /
02	11	uranium	728 N	1127	2010	0.10	10	1 20	509	+6215
32	0		230.0	1000	2010	0.12	13	1.30	550	10,3,4,3
23	V	vanadium	50.94	1890	3380	0.49	b.11	1.63	651	+5,2,3,4
54	Xe	xenon	131.3	-111.8	-107.1	0.158	0.00589	2.6	1170	0
70	Yb	ytterbium	173.1	819	1196	0.15	6.97	-	603	+3,2
39	Y	yttrium	88.91	1523	3337	0.3	4.47	1.22	600	+3
30	Zn	zinc	65.38	419.6	906	0.39	7.13	1.65	906	+2
40	7r	zirconium	91 22	1852	4377	0.27	6 5 1	1.33	640	+4

Table AA. Bond Dissociation Energies & Bond Lengths

Values given are *homolytic* bond dissociation energies, meaning that the electrons are divided equally between the two atoms.

Hydrogen Compounds					
Bond	D	r			
	(kJ/mol)	(pm)			
H – H	432	74			
H – B	389	119			
H – C	411	109			
H – Si	318	148			
H – Ge	288	153			
H – Sn	251	170			
H – N	386	101			
H – P	322	144			
H – As	247	152			
H – O	459	96			
H – S	363	134			
H – Se	276	146			
H – Te	238	170			
H – F	565	92			
H – Cl	428	127			
H – Br	362	141			
H – I	295	161			

Group VIIA Compounds D

(kJ/mol)

155

240

190

148

116

201

273

208

175

Group IIIA Compounds D

(kJ/mol)

293

536

613 456

377

Bond

F – F

CI – CI

Br – Br 1 – 1

At – At

I – O

I – F

I – Cl

I – Br

Bond

B – B

B – O

B – F

B – Cl B – Br r

(pm)

142

199

228

267

?

?

191

232

?

r

(pm)

?

?

?

175

?

Group	Group IVA Compounds						
Pond	D	r					
Boliu	(kJ/mol)	(pm)					
C – C	346	154					
C = C	602	134					
C ≡ C	835	120					
C – Si	318	185					
C – Ge	238	195					
C – Sn	192	216					
C – Pb	130	230					
C – N	305	147					
C = N	615	129					
C ≡ N	887	116					
C – P	264	184					
C – O	358	143					
C = 0	799	120					
C≡O	1072	113					
С – В	356	?					
C – S	272	182					
C = S	573	160					
C – F	485	135					
C – Cl	327	177					
C – Br	285	194					
C – I	213	214					
Si – Si	222	233					
Si – N	355	?					
Si – O	452	163					
Si – S	293	200					
Si – F	565	160					
Si – Cl	381	202					
Si – Br	310	215					
Si – I	234	243					
Ge – Ge	188	241					
Ge – N	257	?					
Ge – F	470	168					
Ge – Cl	349	210					
Ge – Br	276	230					
Ge – I	212	?					
Sn – F	414	?					
Sn – Cl	323	233					
Sn – Br	273	250					
Sn – I	205	270					
Pb – F	331	?					
Pb – Cl	243	242					
Pb – Br	201	?					
Pb – I	142	279					

Group VA Compounds						
Dand	D	r				
Bona	(kJ/mol)	(pm)				
N – N	167	145				
N = N	418	125				
N ≡ N	942	110				
N – O	201	140				
N = 0	607	121				
N – F	283	136				
N – Cl	313	175				
P – P	201	221				
P – O	335	163				
P = 0	544	150				
P = S	335	186				
P – F	490	154				
P – Cl	326	203				
P – Br	264	?				
P – I	184	?				
As – As	146	243				
As – O	301	178				
As – F	484	171				
As – Cl	322	216				
As – Br	458	233				
As – I	200	254				
Sb – Sb	121	?				
Sb – F	440	?				
Sb – Cl ₍₅₎	248	?				
Sb – Cl ₍₃₎	315	232				

Group VIA Compounds						
Dond	D	r				
вопа	(kJ/mol)	(pm)				
0 - 0	142	148				
O = 0	494	121				
0 – F	190	142				
S = O	522	143				
S – S ₍₈)	226	205				
S = S	425	149				
S – F	284	156				
S – Cl	255	207				
Se – Se	172	?				
Se = Se	272	215				

Table BB. Thermodynamic Data

Standard enthalpy of formation (ΔH_f°) & standard entropy (S°) for selected compounds. Note that standard enthalpy values are in kilojoules per mole, whereas entropy values are in joules per mole-Kelvin.

Culture	C 1-1-1	ΔH_f°	S°		Cultert	C 1-1-1	ΔH_f°	S°	Cultert	Charles .	ΔH_f°	S°
Subst.	State	$\left(\frac{kJ}{mol}\right)$	$\left(\frac{J}{\text{mol}\cdot K}\right)$		Subst.	State	$\left(\frac{kJ}{mol}\right)$	$\left(\frac{J}{\text{mol}\cdot K}\right)$	Subst.	State	$\left(\frac{kJ}{mol}\right)$	$\left(\frac{J}{\text{mol}\cdot K}\right)$
Ag	S	0	42.6	ĺ	Cr_2O_3	g	-1139.7	81.2	MnO ₂	S	-520.0	53.0
Ag+	aq	105.79	72.7		Cu	s	0	33.2	N ₂	g	0	191.5
AgCl	S	-127.01	96.2		Cu⁺	aq	+71.7	40.6	NH ₃	g	-45.94	192.3
AgBr	S	-100.4	107.1		Cu ⁺²	aq	+64.8	-99.6	NH_4^+	aq	-133.26	113.4
AgNO ₃	5	-124.4	140.9		CuO	5	-157.3	42.6	NO ₂ -	aq	-104.6	123.0
Al	S	0	28.3	ļ	Cu ₂ O	S	-168.6	93.1	NO ₃ ⁻	aq	-206.85	146.4
Al ⁺³	aq	-538.4	-321.7		Cu(OH) ₂	S	-450	108	N_2H_4	ℓ	+50.6	121.2
AICI ₃	S	-704	110.7		CuS	S	-53.1	66.5	NH ₄ Cl	S	-314.4	94.6
Al ₂ O ₃	S	-1675.7	50.9	l	Cu ₂ S	S	-79.5	120.9	NH ₄ NO ₃	S	-365.6	151.1
Al(OH)₃	S	-1277			CuSO ₄	S	-771.4	107.6	NO	g	+90.2	210.7
Ва	5	0	62.8		F⁻	aq	-335.35	-13.8	NO ₂	g	+33.2	240.0
BaCl ₂	S	-858.6	123.7		F ₂	g	0	202.7	N ₂ O	g	+82	220
BaCO ₃	S	-1216.3	112.1		Fe	S	0	27.3	N ₂ O ₄	g	+9.2	304.2
Ba(NO ₃) ₂	S	-992	214		Fe(OH) ₃	S	-823.0	106.7	Na	S	0	51.2
BaO	S	-553.5	/0.4		FeO	S	-2/2	61	Na ⁺	aq	-240.34	59.0
Ba(OH)₂	S	-998.2	112		Fe ₂ O ₃	S	-824.2	87.4	Na ₂ CO ₃	S	-1131	136
BaSO ₄	S	-14/3.2	132.2		Fe ₃ O ₄	S	-1118.4	146.4		S	-948	102
Be	S	0	10	ł	FeSO ₄	S	-929	121	Naci	S	-411.2	72.1
BeO	S Ø	-599	14		H ₂	g	0	130.6	NaF	S	-5/3.6	51.5
Br ₂ Dr	l aa	121	152.2			uq	26.20	0.0 109.6		S	-407	110
BI-	uy	-121	82 F 7	ł		y aa	-30.29	198.0		S	425.0	04.5
	S P	U 12E 4	5.7			aq	-089.93	91.Z	INI NICL	S	216	29.9
	l P	-135.4 124 E	210.4			y a	-92.31	100.0		S	-310	20 0
	i õ	-134.5	201.7			g ~	-2/3.30	206 5		5	-239.7	38.0
	y a	-74.8 ±226.7	200.8			y aa	20.50	200.5		aq	-230.015	205.0
C2H2	y a	+52.3	200.8			aq	_1200 0	_33.5	D.	y s	0	205.0 164 4
	g	-84.7	210.5	•		aq	-886.0	121.9		3	-287.0	211 7
	g a	-103.8	269.9		H ₂ O	l l	-285 830	69.9		g a	-374.9	364 5
	g f	-238.7	126.8		H ₂ O	a	-241 826	188 7		g a	+5	210
C ₂ H ₂ OH	l	-277.7	160.7		H ₂ PO ₄ -	g aa	-1302.6	90.4	PO ₄ -3	aa	-1277.4	-222
C6H12O6	s	-1275	212		H ₂ N O ₄	a	-20.6	205.7	Pb	s	0	64.8
CO	a	-110.53	197.6		Hg	l l	0	76.0	Pb ⁺²	aa	0.92	10.5
CO2	a	-393.51	213.6	Ì	Hg ⁺²	aa	170.21	-32.2	PbBr ₂	S	-278.7	161.5
CO3 ⁻²	aq	-675.23	-56.9		HgO	cr	-90.79	70.3	PbCl ₂	s	-359.4	136.0
Ca	S	0	41.4		l-	aq	-56.78	111.3	PbO	S	-219.0	66.5
Ca ⁺²	aq	-543.0	-53.1	Ì	12	s	0	116.1	PbO ₂	s	-277.4	68.6
CaCl ₂	s	-795.8	104.6		ĸ	s	0	64.2	PbS	5	-100	91
CaCO₃	s	-1206.9	92.9		K+	aq	-252.14	102.5	PbSO ₄	5	-920	149
CaO	S	-634.92	39.8	1	KBr	S	-393.8	95.9	S	S	0	31.8
Ca(OH)₂	s	-986.1	83.4		KCI	s	-436.7	82.6	SO ₂	g	-296.81	248.1
$Ca_3(PO_4)_2$	s	-4126	241		KClO ₃	s	-397.7	143.1	SO₃	g	-395.7	256.7
CaSO ₄	S	-1434.1	106.7	Ì	KClO ₄	S	-432.8	151.0	SO4-5	aq	-909.34	20.1
Cd	s	0	51.8		KNO₃	s	-494.6	133.0	S	-	2	aq
Cd+2	aq	-75.92	-73.2		Mg	s	0	32.7	Si	5	0	18.8
CdCl ₂	S	-391.5	115.3		Mg ⁺²	aq	-467.0	-138.1	SiO ₂	S	-910.7	41.8
CdO	S	-258.35	54.8		MgCl ₂	S	-641.3	89.6	Sn	S	0	51.6
Cd(OH) ₂	S	-561	96		MgCO ₃	S	-1095.8	65.7	Sn ⁺²	aq	-8.9	-17.4
CdS	S	-162	65		MgO	S	-601.60	26.9	SnO ₂	S	-577.63	52.3
CdSO ₄	S	-935	123		Mg(OH) ₂	S	-924.5	63.2	Zn	s	0	41.6
Cl ₂	g	0	223.0		MgSO ₄	S	-1284.9	91.6	Zn ⁺²	aq	-153.39	-112.1
CI-	aq	-167.080	56.5		Mn	S	0	32.0	Znl ₂	5	-208.0	161.1
CIO ₄ -	aq	-128.10	182.0		Mn ⁺²	aq	-220.8	-73.6	ZnO	S	-350.46	43.6
Cr	S	0	23.8		MnO	S	-385.2	59.7	ZnS	S	-206.0	57.7

					C
Table CC.	some common a	s Equivalent	Units and A	pproximate	conversions

Some Com	Some Common & Equivalent Units						
Length	1 in. (inch)	=	2.54 cm				
	12 in.	=	1 ft. (foot)				
	3 ft.	=	1 yd. (yard)				
	5,280 ft.	=	1 mi. (mile)	=	1,760 yd		
Mass	1 lb. (pound)	=	16 oz.	~	454 g		
	1 ton	=	2000 lb.				
	1 tonne	=	1000 kg				
Volume	1 pinch	=	≤ ¼ teaspoon				
	3 teaspoons	=	1 tablespoon (Tbsp)				
	2 tablespoons	=	1 ounce				
	8 oz. (ounces)	=	1 cup				
	2 cups	=	1 pint				
	2 pints	=	1 quart				
	4 quarts	=	1 gallon				

Some APPROXIMATE Conversions							
Length	1 cm	2	width of a small paper clip				
	6 in.	~	 length of a (US) dollar bill 30 cm 				
	1 ft.	~					
	1 m	~	1 yd.				
	1 mi.	~	1.6 km				
	0.6 km	~	1 mi.				
Volume	1 pinch	~	≤ ¹ ⁄ ⁸ teaspoon				
	1 mL	~	10 drops				
	1 teaspoon (tsp)	~	5 mL	~	60 drops		
	1 tablespoon (Tbsp)	=	3 tsp	~	15 mL		
	2 Tbsp.	=	1 fl. oz.	~	30 mL		
	1 C (cup)	=	8 fl. oz.	~	250 mL		
	1 qt. (quart)	~	1 L				
Mass	1 small paper clip	~	1 gram (g)				
	1 nickel (5¢ coin)	~	5 g				
	1 oz.	~	30 g				
	1 pound (lb.)	=	16 oz.	~	0.5 kg		
	1 ton	=	2000 lb.	~	1 tonne		
Speed	60 mi./h	~	100 km/h	~	30 m/s		
Density	air	~	1 g/L				
	fresh water	~	1 g/mL	~	8.3 lb./gal.		
		~	1 tonne/m ³	~	1 ton/yd. ³		

Table DD. Selected Formulas and Equations							
Density	$D = \frac{M}{V}$	D = density	M = mass	<i>V</i> = volume			
Mala Conversions	1 mol = [molar mass] g	(molar ma	ass = formul	a weight = gram formula mass)			
whole conversions	1 mol = 6.022×10^{23} molecules, atoms or particles						
	measured value - accepted value						
Percent Error $\% \operatorname{error} = \frac{\operatorname{measured value} - \operatorname{accepted value}}{\operatorname{accepted value}} *100\%$							
Percent	% composition - mass o	of part *100%					
Composition	mass of	whole					
	molarity (M) = $\frac{moles of solu}{moles}$	te	normality (N)	= (moles of solute)(dissociation factor)			
	liter of solution	on		liter of solution			
Concentration	molality(m) = $\frac{moles of solution}{kg of solver}$	nt					
	mole fraction $(\chi_A) = \frac{\text{moles of}}{\text{total mo}}$	F A les	parts per mi	$ lion(ppm) = \frac{grams of solute}{grams of solvent} \times 1,000,000$			
	$P_1 V_1 = P_2 V_2$	P = pressure		n = moles			
Casas	$\frac{1}{T_1} = \frac{2}{T_2}$	V = volume(1)		T = temperature (K)			
Gases		P = 0.0021 Liat	m o o a Liki	$P_{a} = c_{a} + t_{b}$			
	PV = nRI	$R = 0.0821 \frac{2 d c}{mol}$	$\frac{1}{K} = 8.31 \frac{1}{mo}$	$\frac{1}{100} = 62.4 \frac{1000}{\text{molK}}$			
	1 atm = 101.3 kPa =	760 torr = 760 n	nm Hg = 29	.92 in. Hg = 1.013 bar = 14.7 psi			
Pressure	$P_{\rm A} = \chi_{\rm A} P_{\rm T}$	P_{A} = partial pressure of A χ_{A} = mole fraction of A					
	$P_{\rm T} = P_{\rm A} + P_{\rm B} + P_{\rm C} + \dots$	$P_{\rm T}$ = total press	ure				
Titration	$N_A V_A = N_B V_B$	N_A = normality V_A = volume of	of H₃O⁺ acid	N_B = normality of OH ⁻ V_B = volume of base			
	$\Delta T_f = f$	reezing point dep	ression	The balling a give allowed in a (20)			
	$\Delta T_f = imK_f \qquad (°C)$		Δ	$T_b = \text{boiling point elevation (°C)}$			
	$K_f = free$	ezing point depre	ession const	ant (H2O = 0.52 °C/m)			
Colligative	$\Delta T_b = imK_b$ $K_b = bc$	oiling point elevat	ion constant	$(H_2O = 1.86 \text{ °C/m})$			
Properties	i = va	't Hoff factor (diss	sociation fac	tor) π = osmotic pressure			
		= molality	M = mola	rity $N = normality$			
	$\pi = IIVI RI = IV RI$	ideal gas const	(mol/L)	(moi/L)			
			r – tempe				
		$K_a = \frac{[H_3 O][A]}{[H_3 O]}$	p <i>K</i> a = −log	$g(K_a)$ [base]			
Acid-Base	pOH = -log[OH ⁻]	[ΠΑ]		$pH = pK_a + \log \frac{[busc]}{[busc]}$			
	рН + рОН = 14	$K_{h} = \frac{[\mathrm{H}^{+}\mathrm{B}][\mathrm{OH}^{-}]}{1}$	$pK_b = -log$	(K_b)			
	$K_w = K_a \cdot K_b = 1 \times 10^{-14}$	5 [B]					
Equilibrium	$\mathcal{K}_{\rho}=\mathcal{K}_{c}(R T)^{\Delta n}$	K _p = gas press. equil. const	<i>K</i> _c =	molar conc. Δn = change in equil. const. # moles			
	$q = m C_n \Delta T$	q = heat	-	ΔT = change in temperature			
Heat	$q = m \Delta H_f$	<i>m</i> = mass		ΔH_f = heat of fusion			
	$q = m \Delta H_v$	C_p = specific hea	t capacity	ΔH_v = heat of vaporization			
- 1 1 ·	$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$	G° = standard fr	ee energy	S° = standard entropy			
Inermodynamics	$\Delta G^{\circ} = -RT \ln K$	H° = standard er	nthalpy	T = temperature (K)			
				<i>q</i> =			
	a	/ = current (amn	eres)	charge $t = time(seconds)$			
Flectrochemistry	$I = \frac{q}{t}$	<i>i</i> = current (unip	ci caj	(Coulomb			
Licenseneniisti y	•			s)			
		F = Faraday's co	nstant = 96,	000 Coulomb per mole electrons			
	$\Delta G^{\circ} = -nFE^{\circ}$	n = moles of ele	ctrons	<i>E</i> ° = standard reduction potential			
Temperature	K = °C + 2/3.15 °C= (°F-32) * 5/9	K = Kelvin	sius	°F = degrees Fahrenheit			
	$\frac{t}{1} = \frac{1}{2} $	$\frac{c - degrees cer}{t = tc}$	ntal elansed	time t			
Radioactive Decay $A = A_0 \left(\frac{1}{2}\right)^{\tau_{\chi_2}}$ $A_0 = \text{original amount}$ $\tau_{\chi_2} = \text{half-life}$ number of half-lives = $\frac{1}{\tau_{\chi_2}}$							

Periodic Table of the Elements



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