CHEM 1300: PREPARATORY CHEMISTRY



Robert Belford
University of Arkansas Little Rock

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Chem 1300: Preparatory Chemistry

Robert Belford

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0: GENERAL INFORMATION

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Learn more

Welcome to your new guide!

Click on "New" to build out your guide.

- Topic hierarchy

0.1: Creating Accounts

0.2: Google Classroom

0.3: Genius Scan- Cell Phone App

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0.1: Creating Accounts

Creating Libretext ADAPT Account

You will use the ADAPT system of LibreText to submit your assignments and you will use the Single Sign On option through your **school (@ualr.edu) email. Do not create a private account within the ADAPT system.**

1. Go to https://adapt.libretexts.org. (Figure 0.1.3) Click **Log IN** (DO NOT REGISTER!)

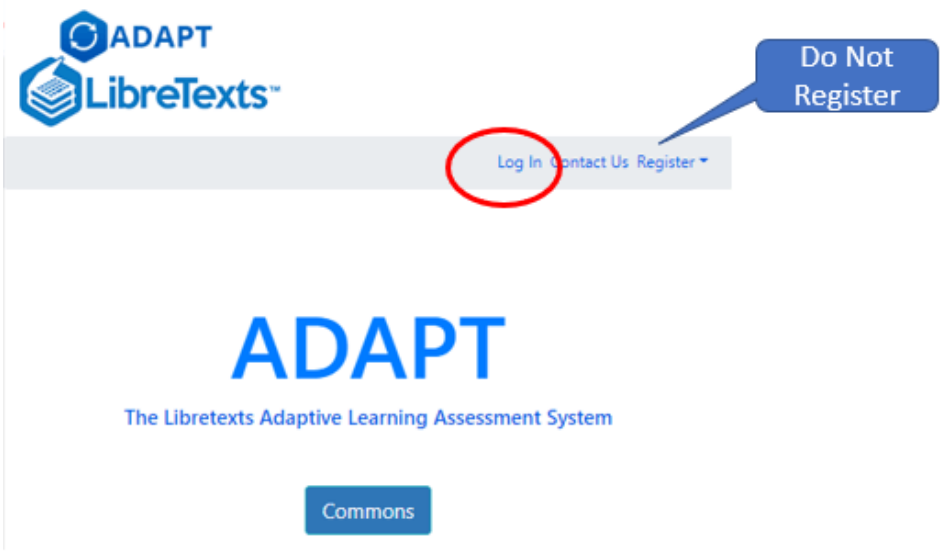


Figure 0.1.3: Click Log In. Do not Register.

2. Click Campus Login **(DO NOT FILL OUT FORM**) (figure 0.1.4).

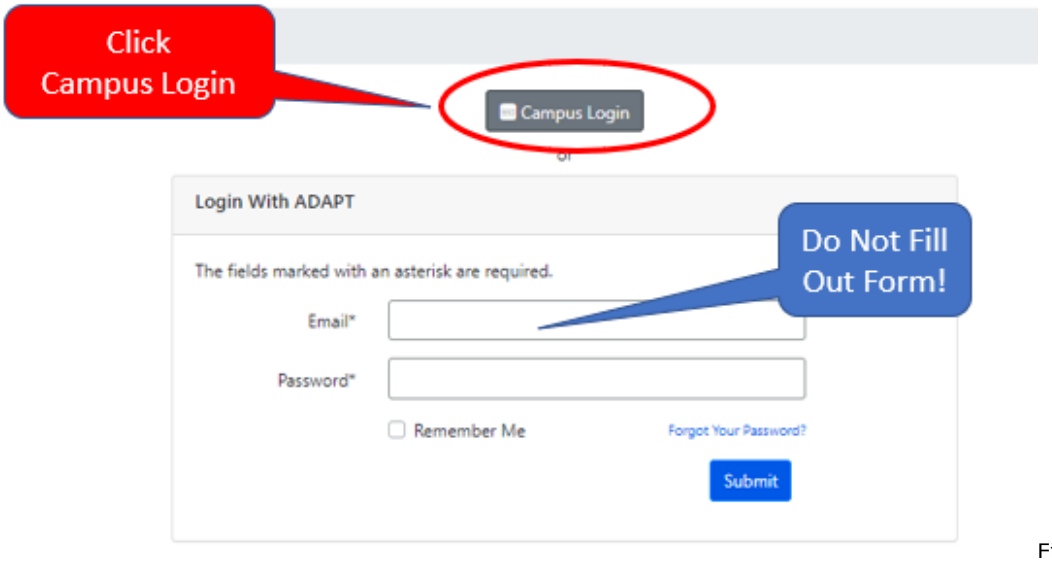


Figure 0.1.4: Click "Campus Login", this will allow you to set up a Single Sign On (SSO) with your UALR email account.

3. Click "Google Workspace (UALR email is run by Google) (figure 0.1.5).

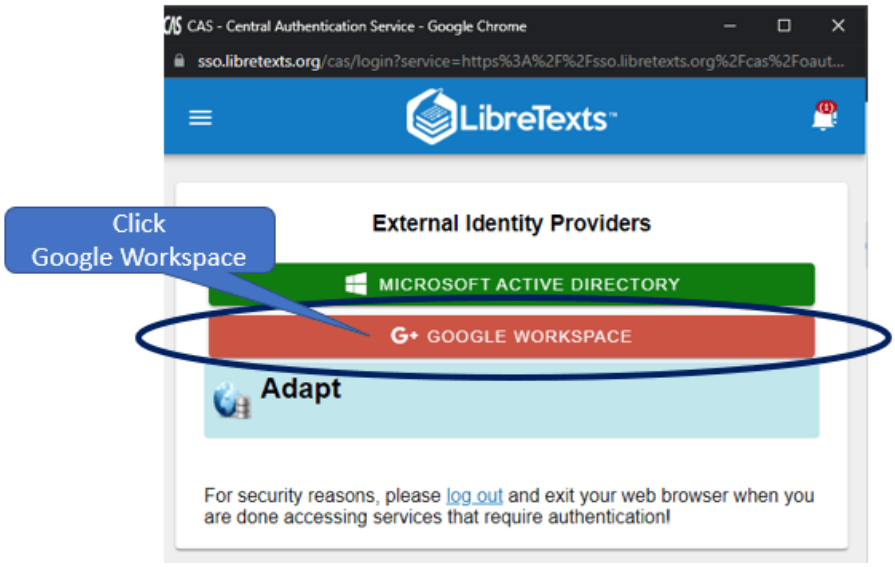


Figure 0.1.5: You will be able to log in with your UALR email through the Google Workspace option)

4. Choose your UALR email account and fill out form, choosing the student option (figure 0.1.6).

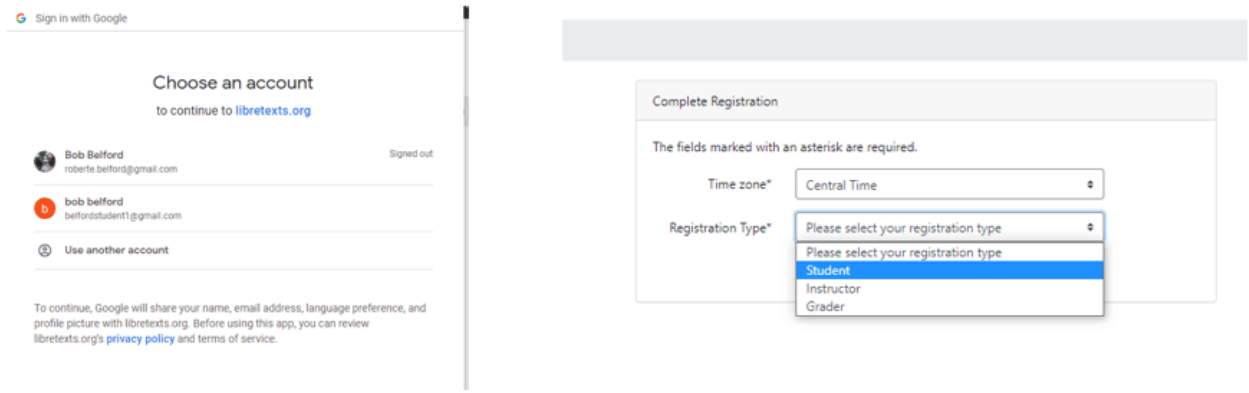


Figure 0.1.6: Be sure to set up the account with your UALR email address.

5. You now have an account and can enroll in a course.

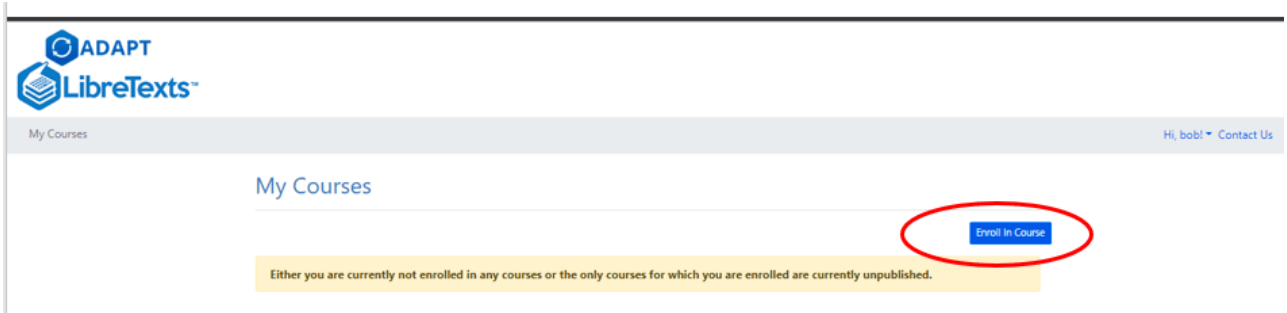


Figure 0.1.7: You now have an account within the ADAPT system and need to get an access code from your instructor to enroll in a course (Copyright; author via source)

6. You will need an access code for each course you take. If you are in a lab, each lab section has its own access code. These will be provided by your instructor.



Figure 0.1.8: Insert your access code and you will become registered for a course.

7. You now can navigate to your course assignments

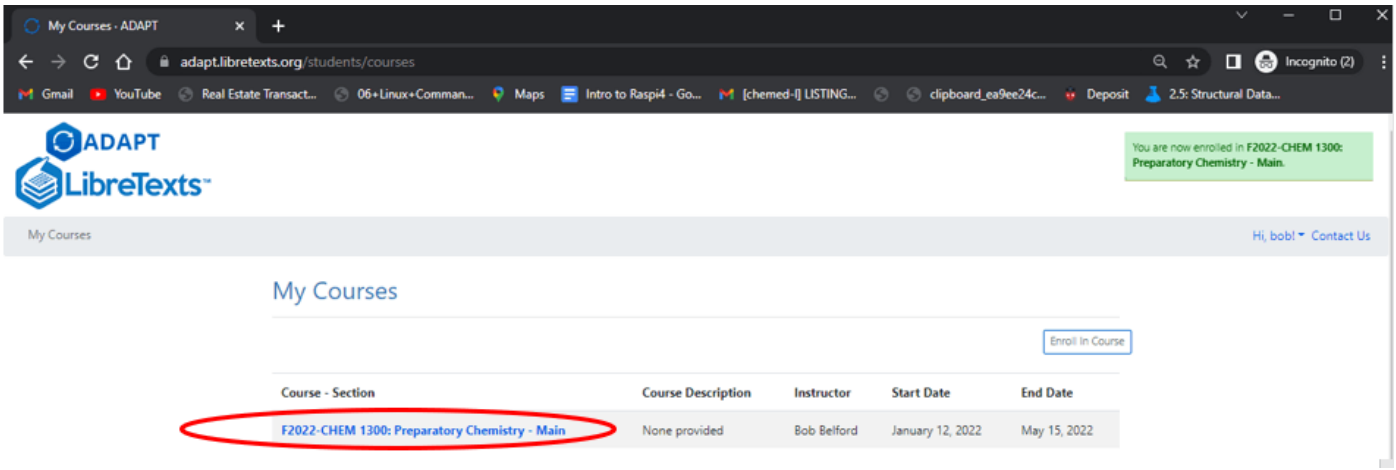


Figure 0.1.9: Copy and Paste Caption here. (Copyright; author via source)

Setting up Notifications

ADAPT can send an alert to your email when an assignment is due. This is advised as you will lost points for late assignments.

1. Next, Click on your name in the top right corner and go to Settings.

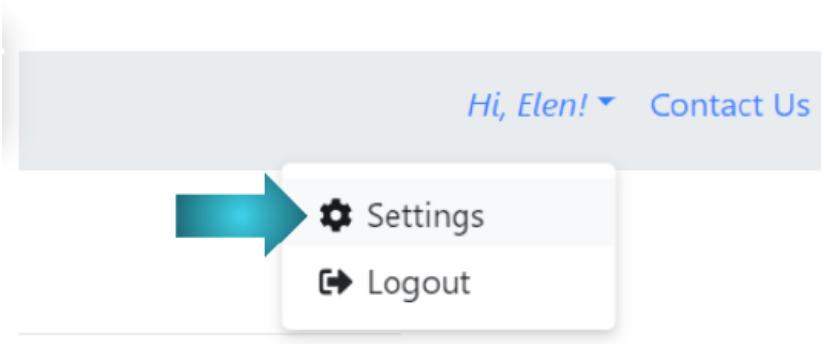


Figure 0.1.10: Open settings

2. Go to Notifications and choose the time option for your reminders. Click "Update"

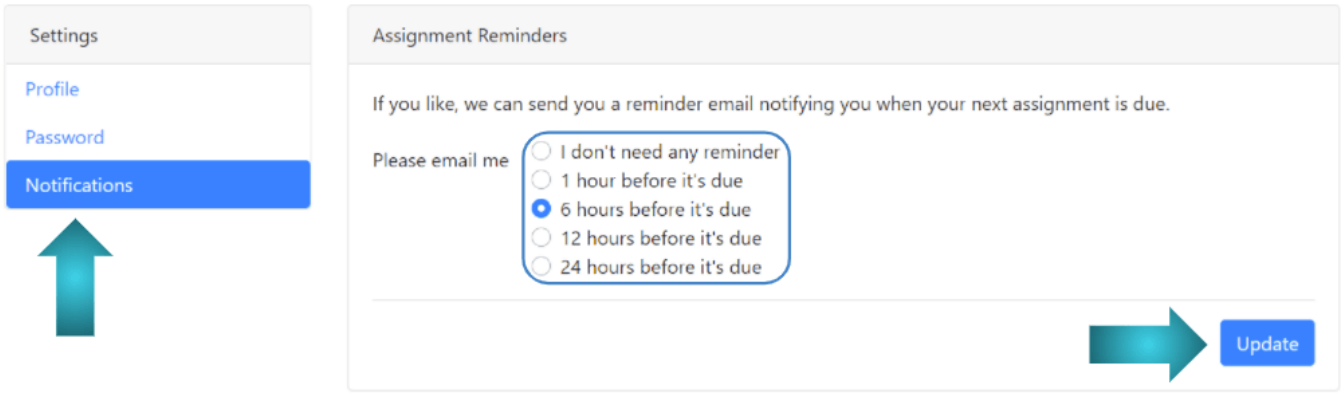


Figure 0.1.11: Copy and Paste Caption here. (Copyright; author via source)

Creating Zoom Account

First thing you need to do is create a Zoom account if you don't have one. It is recommended that you use your school (@ualr.edu) email to create your Zoom account. If you already have an account, please make sure that your name is easily identifiable - first and last name is ideal.

1. Navigate to https://zoom.us/
2. Click on the big orange button in the upper right corner that says "Sign Up, it's Free"

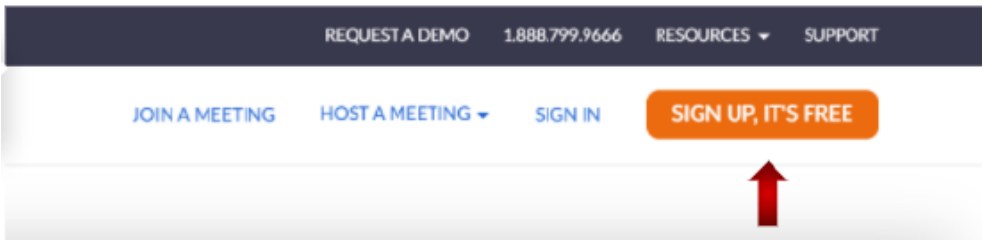


Figure 0.1.1: Click on the button indicated in the image

3. Input your date of birth. This information is used for verification and will not be stored.
4. On the next page, input your email (it is recommended that you use your @ualr.edu email). You can also sign up with your Google account and select your school email.
5. Go to your email and look for the activation email. Open it and click “Activate account”. If you don't see the email, please check your spam folder, social and promotional tabs (if you use Gmail).

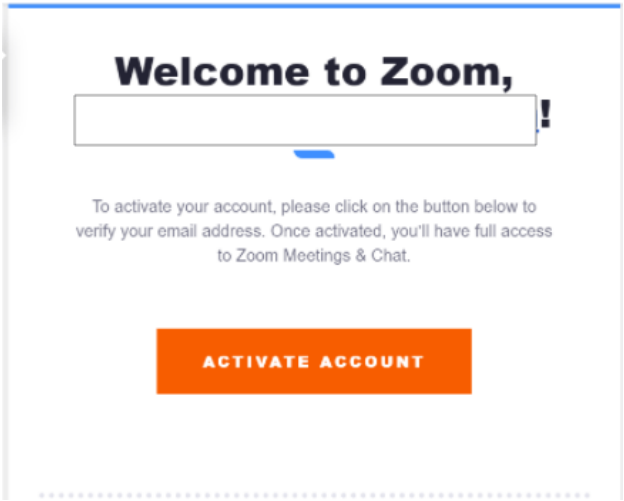


Figure 0.1.2: Activation email from Zoom

6. Fill out your first and last name and create a password. Make sure to use a strong password.
7. Let Zoom know that you are not a robot by completing reCAPTCHA.

8. Skip the next step and then navigate to your account.

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0.2: Google Classroom

Accessing Google Classroom

If you open your Google Mail you will see 9 dots by your name which opens the Google Suite. If you scroll to the bottom you will find Google Classroom, and if you open the Drive you will also find a folder called Classroom where you can find your assignments.

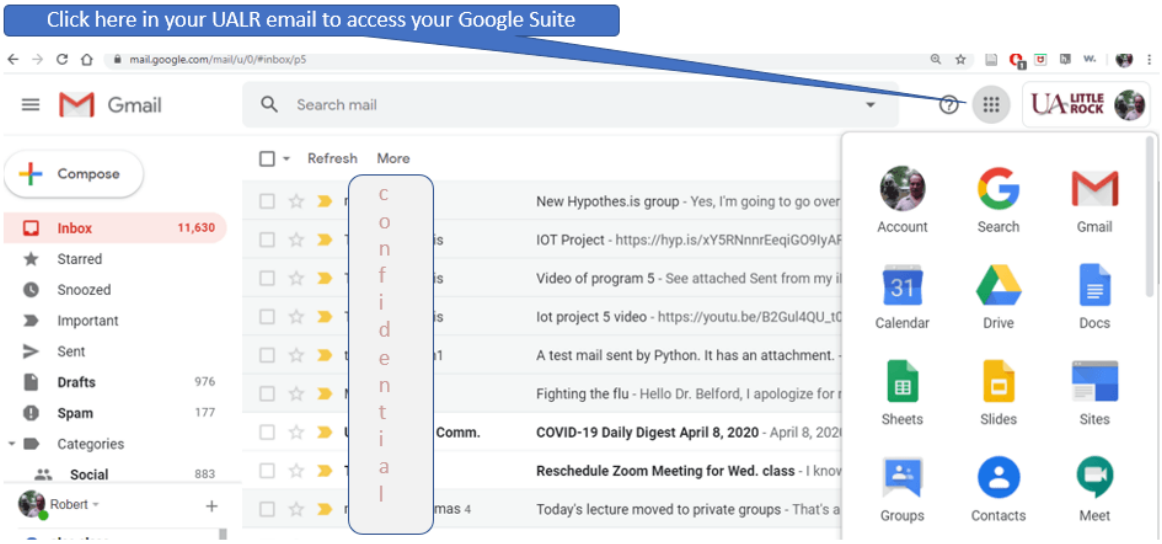


Figure 0.2.1: How to access the Google suite.

Email Notifications

This class will use Zoom and you need to be logged in by 9:00 AM. I suggest you set up an email notification through your Google Calendar, and the following video shows you how to do that for the duration of the class (with the email containing the link being sent 10 minutes before class).

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

The above video was created for the Summer 2020 general chemistry 1 class and shows how to set up notification with Google Calendar. In your notification you can include links, and this can be a convienent way to get the class link emailed to you. Our goal is to run this class in person, but in the event of a COVID surge you need to be prepared to use online resources like this.

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0.3: Genius Scan- Cell Phone App

Genius Scan

Some of your Daily Quizzes will require you to upload a PDF. There's a quick and easy way to do that if you don't have access to a scanner.

Step 1 Set up email app on your mobile device. There are instructions available for iOS devices and Android devices.

Step 2 Download Genius Scan. You can use links given below if you are viewing this page from your phone, or you can simply search for this app. Free version is enough.

- Apple App Store
- Google Play

Step 3 Open Genius Scan and add scan your first document.

1. Open the app.
2. Click on the plus sign in the bottom left corner.
3. Allow access to the Camera.
4. Position your phone over the sheet of paper you want to scan. **The whole page has to be visible in your camera view.** Click "Scan" (big round button at the bottom). You can scan as many pages as you need, Genius scan will save them all in one PDF document, and you would be able to slice it in ADAPT.
5. Once you've scanned all pages hit "Done".

Please keep all scanned pages as is. **DO NOT crop your PDF files** as it will result in a loss of part of the page during upload to ADAPT.

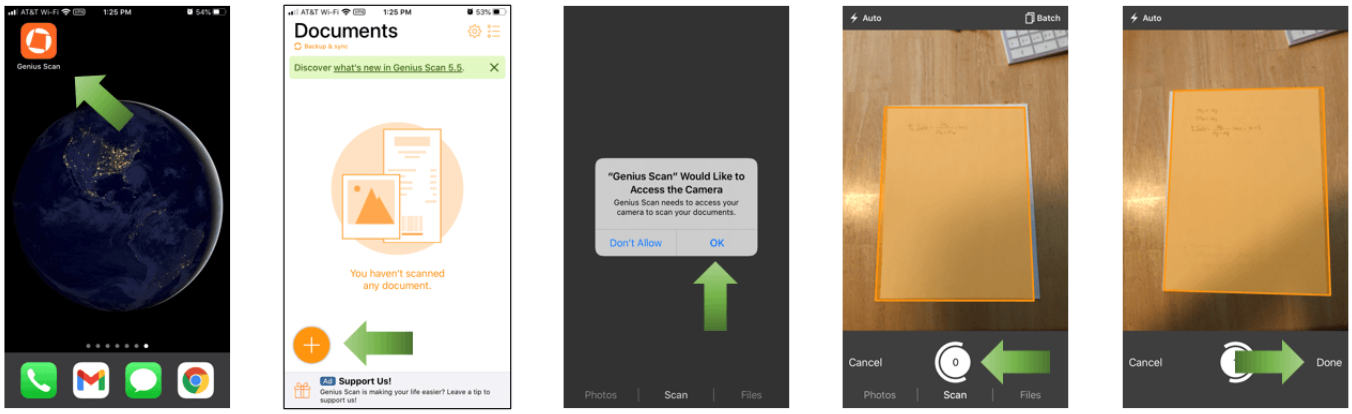


Figure 0.3.1: Scan as many pages as you need. NOTE the yellow outline of the page in images 4 and 5, that is indicating that the scan will go to a separate page. Do not have the answer to two questions on the same page, as they will be uploaded to different sections of the ADAPT submission system.

Step 4 Email your PDF file to yourself and access it from your computer to upload as a file submission. Refer to Figure 0.3.2 for instructions.

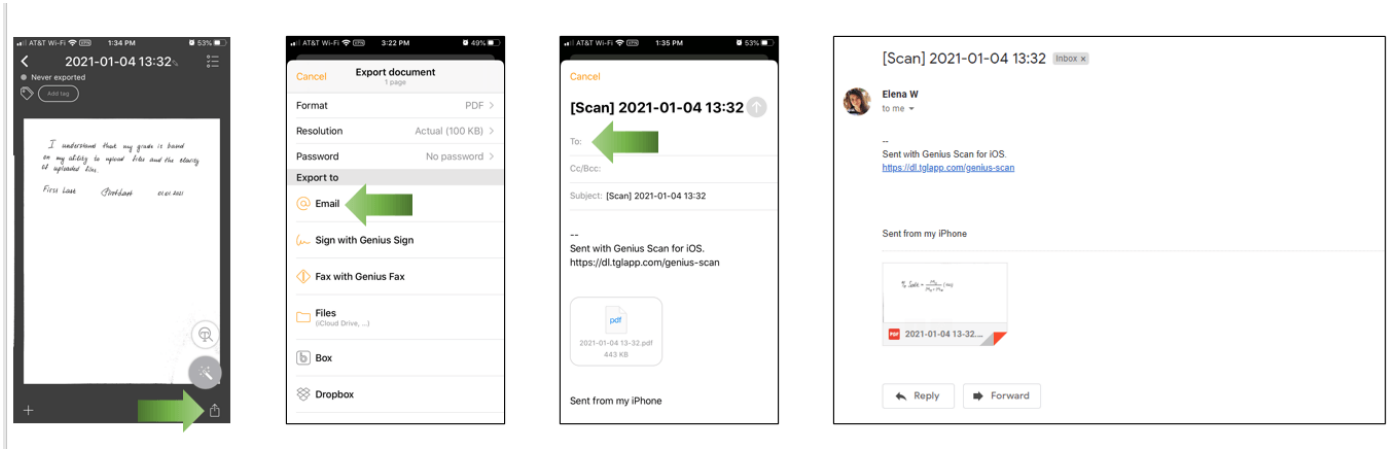


Figure 0.3.2: Email your scan

Notes on Taking Scans

To achieve the best results, before uploading the file, make sure that it is:

- Clear and presentable
- Easy to read and follow
- Centered and the entire solution is visible

You can see an example of an acceptable submission in Figure 0.3.3

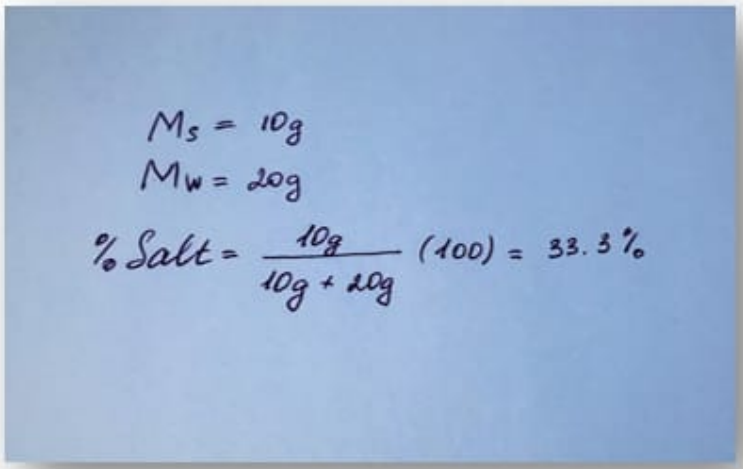


Figure 0.3.3: This image is clear, easy to read, and the entire solution is visible (you will also need to write your name, signature, and date on the sheet of paper)

Below are some examples of **unacceptable** submissions. You will lose points if the work is not clear.

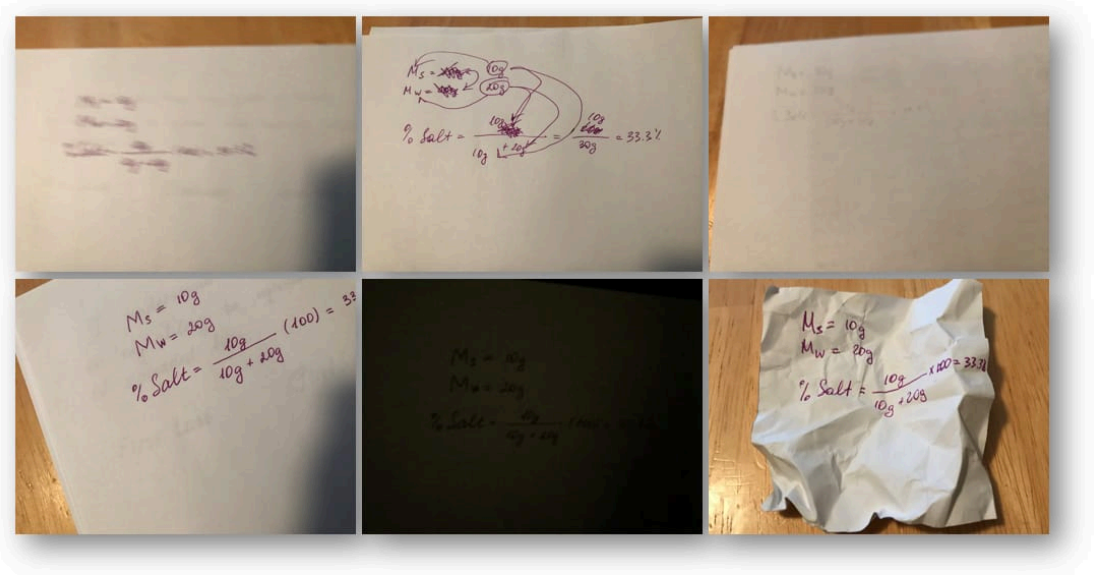


Figure 0.3.4: Unacceptable submissions

It is advised to take multiple scans of each answer, then when you upload them you can choose the ones you want to assign to each question.

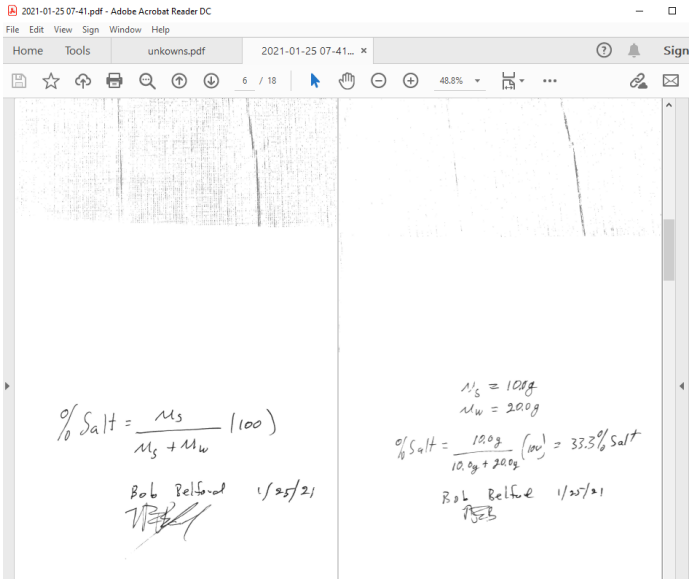


Figure 0.3.5: Note the above PDF has 18 pages with a scan on each page. This allows you to pick which image you want to submit as the answer to each question.

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1: INTRODUCTION

We are in the middle of the COVID19 pandemic and this introduction provides two types of information content. The first deals with the topic of the class, and the second deals with the technologies used to run the class in a hybrid environment.

- Topic hierarchy

1.1: Chemistry as a Science

1.2: Mathematics and Scientific Communication

1.3: Chemical Safety

1.4: Using ADAPT

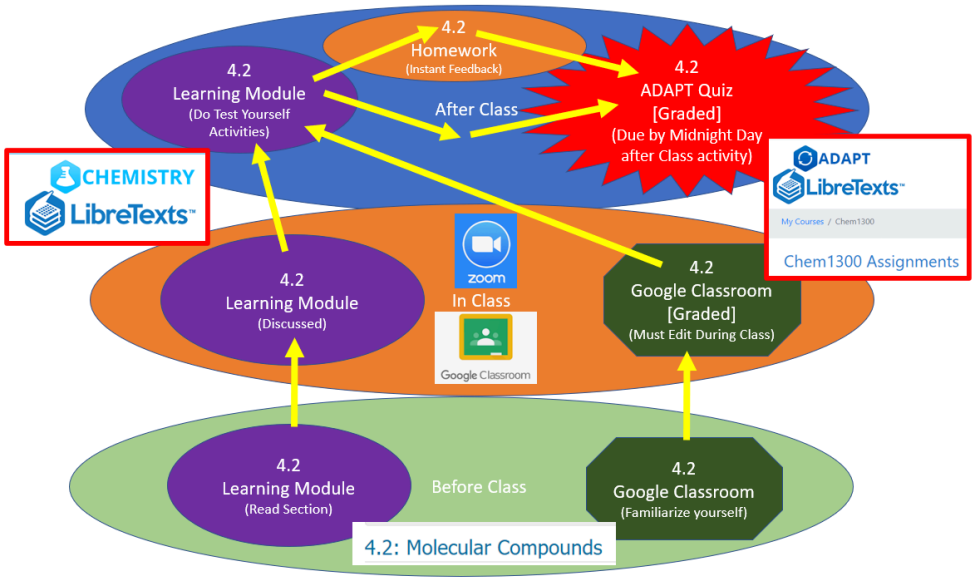


Figure 1.1:Suggested diagram on how to navigate the hybrid components of the class.

Many student find taking a hybrid class to be more challenging than a regular class and simply speaking, we need to use digital technologies. We will be using multiple digital technologies and the sooner you master them the better off you will be. We are using the same table of contents across all platforms, so section 4.2 Deals with Molecular Compounds, and the assignments are all related. Figure 1.1 outlines a prudent schema for traversing the course material. It is suggested that before each lecture you read the LibreText part on that lecture and review the Google Classroom activity. We will typically spend the first half hour going over material, and the remaining time working on the Google Doc in Google Classroom in small groups using Zoom Breakout rooms. You need to edit the Google doc during the class period, not before or after, and the Google Classroom portion of your grade will be based on whether you show up in the history of the document. As soon as the class is over you will be provided the key, and you have until midnight of the next day to take the quiz in ADAPT. But you only get one chance at the quiz, and it is suggested that you review the "Test Yourself" section of the LibreText learning module, and if necessary, the associated homework. These all give instant feedback, and if you need more help, go to the online Tutor sessions, as posted in Google Classroom. Once you think you have mastered the material, go to that section in ADAPT and take the quiz. Many of the quizzes will require you to take a photo of your work and upload it to ADAPT.

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1.1: Chemistry as a Science

What is Chemistry

Let's start with Webster's Dictionary:

> *a science that deals with the composition, structure, and properties of substances and with the transformations that they undergo (*"Chemistry." *Merriam-Webster.com*. Merriam-Webster, n.d. Web. 16 May 2016.)

And Oxford's Dictionary:

> *The branch of science that deals with the identification of the substances of which matter is composed; the investigation of their properties and the ways in which they interact, combine, and change; and the use of these processes to form new substances. (*chemistry. Oxford Dictionaries. Oxford University Press, n.d. Web. 17 May 2016.)

So Chemistry is a Science.

What Distinguishes a Science From Other Types of Academic Disciplines?

This is a good question, and I would say the **Scientific Method**.

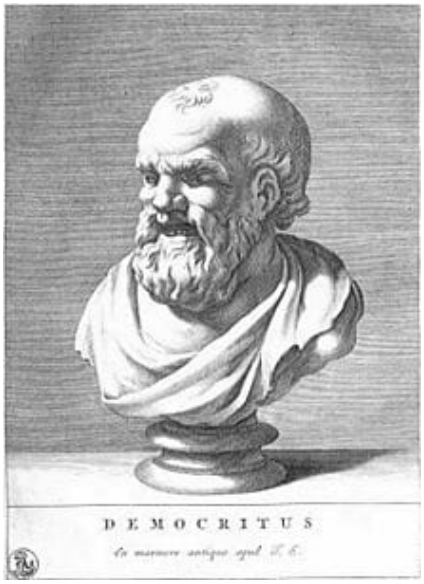




Figure 1.1.1: Democritus (right, public domain) Greek amphitheatre, Siracusa (right, cc-by-2.0 Andrea Schaffer)

Democritus of pre-socratic Greece came up with an ancient atomistic theory that viewed mater as being based on two fundamental entities, atoms and the void. This ancient theory was similar in principle to the modern atomistic theory and is in fact where the name "atom" came from. The fundamental difference is that it was not based on the Scientific Method and the rigorous verification based on empirical (measurable) data. In essence, many philosophies could produce understanding and frameworks for understanding that gained societal acceptance through oration and debate, but just because they were widely accepted, did not mean they were correct. The scientific method places faith on measurable entities and the repeatability of the causal relations between them.

So What is the Scientific Method?

The Scientific method is a philosophy based on the process of making observations, forming hypothesis that could explain the observations, and testing the hypothesis through experimentation.

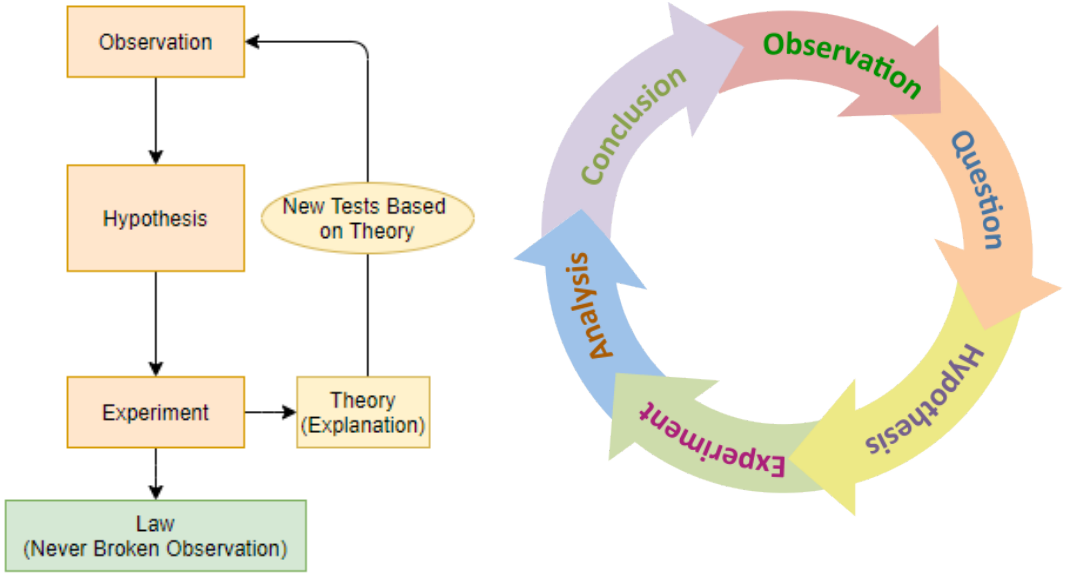


Figure 1.1.1: Two flowcharts of the Scientific Method (left by Belford, right cc.1⓪ by Thebiologyprimer). The flow chart on the left shows the role of theory and how scientific laws can evolve while the one on the right emphasizes the continual evolving nature of scientific knowledge.

A **theory** is a mental construct or model that explains what is happening, and is often the basis for further hypothesis to test with additional experiments. A theory can be right for the wrong reason. A scientific **law** is a refined result of this process that can never be broken and allows us to predict what will happen, but does not explain why. For example, the first law of thermodynamics says that the energy of the universe is constant, so if heat is transferred from one object to another in an isolated system, the heat lost by one will be the same quantity as the heat gained by the other. If a process occurs where energy is not conserved, the first law of thermodynamics would be invalidated, and it would no longer be a law.

The Scientific Method is sort of the philosophy of which scientific knowledge is based upon

---

What is the difference between Qualitative and Quantitative Information?

Qualitative information is based on observables, like ice floats on water, or chlorophyll is green, or the patterns of the lines on your fingerprint. Qualitative data can be used to identify what something is.

Quantitative information relates to a measurable quantity. Quantitative data may be a number like the number of seats in a room, but it also may have a number and a unit, like the number of grams a dime weighs. Quantitative data tells you how much of an observable you have.

How Old is Chemistry as an Applied Science?

As an applied science, chemistry dates to antiquity.

File:Primitive man making tools and using fire at the Museum of Vietnamese History.JPG



Figure 1.1.3: Primitive man making tools and using fire, (left, cc 3.0 - Museum of Vietnamese History) and possibly 52,000 year old painting of bovine from Lubang Jeriji Saléh cave (right, public domain)

Primitive man learned how to make and keep fire that allowed him to cook and keep warm. Primitive man learned how to extract pigments from minerals that were used to paint their caves, and some of these have lasted multiple thousands of years. There were often religious overtures to their use of applied chemistry as shamans used these technologies to interact with their spirit world.

How old is Chemistry as a Modern Science?

Let's compare the text on two documents, one from 1776 and one from 1808.

File:US Declaration of IndependenceFile:Daltons symbols.gif
us0036 03.jpg

Figure 1.1.4: On the left is an image of the U.S. Declaration of Independence written in 1776, and on the right an image from John Dalton's 1808 book, "A New System of Chemical Philosphy" (both images in Public Domain).

The image on the left is the U.S. Declaration of Independence, while the image on the right comes from John Dalton's 1808 book, "A New System of Chemical Philosophy", which was written 32 years after the United States Declaration of Independence. Note that even in these recent times, hydrogen was not "H" but a circle with a dot in the middle of it, and Oxygen was a circle, and water was the two next to each other. The symbols of the periodic table that are so common today did not even exist when George Washington was president!

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1.2: Mathematics and Scientific Communication

Introduction

Before covering any chemistry it is important to understand the difference between algebra and arithmetic, and how they are applied to scientific theory. In this activity we will jump to material that will be covered later in the semester, like the First Law of Thermodynamics, and our goal is not to solve a thermodynamic problem, but to understand the difference between an algebraic equation and an arithmetic calculation. That is, to the scientist algebra is a language that can state knowledge through the use of variables that have physical meaning and significance, and just like a sentence in the English language, it can be read and understood by someone who knows the science. The algebraic equation (sentence) can mathematically be solved for a specific variable whose value can be arithmetically determined if the values of the other variables are known, but once you substitute that numerical values into the variables, it can no longer be read like a sentence in the English language. Figure 1.2.1 shows a flow chart that relates these where you develop an algebraic equation based on scientific theory and law, solve that for the value you seek and then substitute in the numbers to produce the answer you desire.

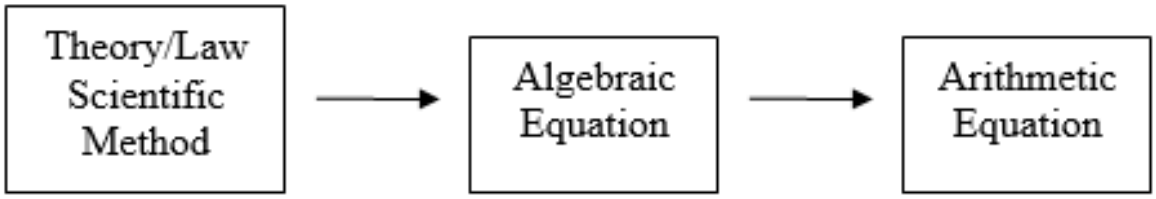


Figure 1.2.1: Flow Chart for applying scientific theory or law to solve a problem (see example 1.2.1 below.

Lets apply Figure 1.2.1 to a scientific problem that we will need to be able to solve before this semester is over. We could ask the following question in English.

Example 1.2.1

What is the final temperature if a 1 g piece of aluminum at 80^0 C is dropped into 1 mL of water at 20^0 C in a thermally isolated system?

Solution

To solve this problem we need to derive an algebraic equation from scientific theory and law. Later in this semester you will learn about the First Law of Thermodynamics, which states that the Energy of the Universe is constant, and that for an isolated system the heat lost by a hot object is the heat gained by a cold object if they are in a thermally isolated system. This can be described by the following algebraic equation, where Q_h is the heat leaving the hot object (note there is a negative sign) and Q_c is the heat gained by the cold object.

$$-Q_h = Q_c$$

But the above equation gives us no information relating this energy flow to the temperature and so as written, can not be used to solve our problem. But as we shall learn later this semester, when you add heat to a substance one of three things typically happens

a. The temperature rises
b. A phase change occurs
c. A chemical reaction occurs

For a system at constant phase where there is no reaction occurring (the iron does not heat the water so much that it boils or breaks into hydrogen and oxygen), we can relate the heat transferred and temperature change to a property of the matter called its heat capacitance, which can be expressed as

$$Q = C\Delta T \tag{1.2.1}$$

or

$$Q = mc\Delta T \tag{1.2.2}$$

where for aluminum c_{Al}= 0.902J/(g-K) and for water c_{Water}= 4.184J/(g-K)

substituting eq into eq gives

$$-m_h c_h(T_F-T_H) = m_c c_c(T_F-T_c)$$

Which can be solved for the final Temperature

$$T_F = \frac{m_H c_H T_H + m_c c_c T_c}{m_H c_H + m_c c_c} \tag{1.2.3}$$

Once this is done, you plug in your numbers and do the arithmetic.

$$T_F = \frac{(1.00g)(0.902\frac{J}{g\cdot^oC})80.0^oC + (1.00ml)(\frac{1.00g}{1.00ml})(4.184\frac{J}{g\cdot^oC})20.0^oC}{(1.00g)(0.902\frac{J}{g\cdot^oC}) + (1.00ml)(\frac{1.00g}{1.00ml})(4.184\frac{J}{g\cdot^oC})} = 30.0^oC \tag{1.2.4}$$

or worse, you could do the math as a pure arithmetic problem

$$T_F = \frac{(1.00)(0.902)80.0 + (1.00)(\frac{1.00}{1.00})(4.184)20.0}{(1.00)(0.902) + (1.00ml)(\frac{1.00}{1.00})(4.184)} = 30.0^oC \tag{1.2.5}$$

The important message is you can read eq. _____ but can not read eq. ____. That is, eq. is an algebraic equation, and a scientist who has passed freshmen chemistry can read the equation and identify if there was a mathematical mistake when it was solved, where as eq. is just an arithmetic operation where you plug the numbers into a calculator and come up with an answer.

Can you see another potential issue with eq.____? There are no units, and you should always include units when you convert you algebraic equations to arithmetic ones, as the units should cancel in a manner where the answer has the same units on both sides of the equations (if you are solving for temperature, your answer needs the units of temperature or you made a mistake.

IMPORTANT: PARTIAL CREDIT
in this class
WILL ONLY BE GIVEN IF THE ALGEBRAIC EQUATIONS ARE SHOWN,
and when a problem states
"SHOW CALCULATIONS"
you must show both the
algebraic and the arithmetic calculations.

This is very important and a very common mistake that students make. They solve the arithmetic problems and not the algebraic, and the sooner you learn to solve the algebra, and then do the arithmetic, the better off you will be. That is the purpose of this example, and anyone who knows chemistry can read eq (), and see if there is a mistake, but eq() can not be read, it is just a bunch of numbers, that could mean anything.

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1.3: Chemical Safety

General Information

Chemical Safety is usually covered in the chemistry laboratory and we will not be doing any experiments in this class. It is still important to discuss safety and go over some resources that are available to the public. One of the goals of this section is to provide you resources for chemical safety in the digital age, when it is so easy to find misinformation. We will also discuss the evaluation of safety information, which is ever so important now, in the time of COVID-19, and this material is sort of separate from the rest of the class, and will not be examined on, but is here for your benefit. So lets start with workplace safety, and the Chemical Hygiene Plan at UALR.

Chemical Hygiene Plan

Every university has a CHP (Chemical Hygiene Plan) as required by OSHA standard 29 CFR 1910.1450 and a CHO (Chemical Hygiene Officer) who is responsible for its implementation, and UALR's CHP can be found at the Facilities Management Web Page. Within the CHP are a set of standard SOPs (Safe Operating Procedures) that represent the minimum safe practices for the handling of hazardous chemicals and operation of equipment (like how to secure a compressed gas tank). Every research lab at the university is required to develop and maintain SOPs for the laboratory practices that are performed within their labs, and here is a link to the template for developing laboratory specific SOPs. The CHP also defines the PPE (Personal Protective Equipment) needed to perform work in a laboratory, and in the university teaching laboratory the instructor would be responsible for ensuring students abide by the established SOPs, and wear proper PPE, like safety glasses and closed toe shoes (no flip-flops in the chemistry laboratory), and do not perform any unauthorized experiments.

Chemical Safety Resources

Prudent Practices

The National Research Council of the National Academies of Sciences has published a book "Prudent Practices in the Laboratory" that can be downloaded for free and has a wealth of information on chemical safety, including a copy of OSHA's Laboratory Standard (29 CFR 1910.1450). There is also an accompanying zip file of a CD that contains Laboratory Chemical Safety Summaries (LCSS) and additional information. LCSS are very useful and in a bit we will look at the PubChem LCSS that are freely available over the web.

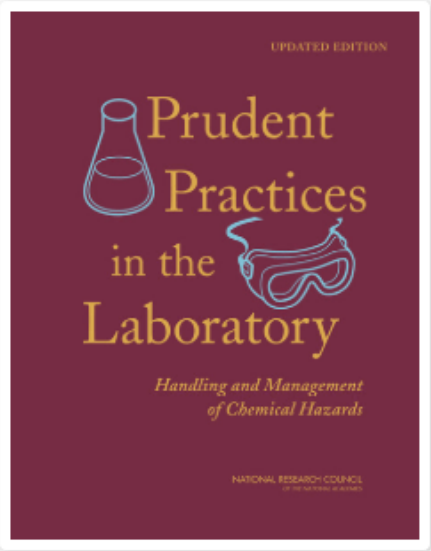
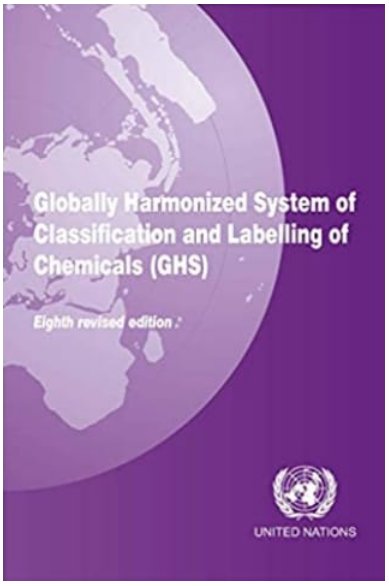


Figure 1.3.1: NRC's Prudent Practices in the Laboratory.

Do you know who is considered the founder of the National Academy of Science? Check it out on Wikipedia! Yes, a president of the United States is considered to be a co-founder of the National Academy of Science and it was created by an Act of Congress to "provide independent, objective advice to the nation on matters related to science and technology...." (see Wikipedia article). But providing information on chemical safety can be very complicated, why before 2012 we had different symbols for chemicals depending on if they were being transported (DOT regulations), used (OSHA regulations) or being disposed of (EPA regulations), and of course, we live in a global world where chemicals are shipped across borders, and so there are not only different laws, but different languages.

UN GHS

What a mess! So the United Nations put forth a system, the Globally Harmonized System of Classification and Labeling of Chemicals (GHS) and the OSHA's laboratory standard is actually integrated into the United States implementation of the UN GHS. The 8th edition can be downloaded as a PDF for free and you can find codes that mean the same thing in any country that adheres to the UN recommendations. So if a chemical is made in China, travels through Indonesia, Saudi Arabia, Mexico and is used in the United States, everyone knows how to read the hazard codes. Another component of the GHS is the Safety Data Sheets (SDS), which provide hazard information outlined in figure 1.3.2 and are required for every chemical that is manufactured or enters the US. So if you want to find information on a product, go find the SDS, and further information bay be obtained at Safety Emporium.



1.5.3.2.1 The information in the SDS should be presented using the following 16 headings in the order given below:

1. Identification
2. Hazard(s) identification
3. Composition/information on ingredients
4. First-aid measures
5. Fire-fighting measures
6. Accidental release measures
7. Handling and storage
8. Exposure controls/personal protection
9. Physical and chemical properties and safety characteristics
10. Stability and reactivity
11. Toxicological information
12. Ecological information
13. Disposal considerations
14. Transport information
15. Regulatory information
16. Other information.

Figure 1.3.2: On left is an image of the 2019, 8th edition of the UN GHS and on right are the 16 GHS mandated categories for chemical SDS.

And interesting thing about the UN GHS is they allow us to compare the laws of different nations. That is, the laws of a country determine what codes must be applied to a chemical, and since they all use the same code, you can quickly compare the laws. So when labeling a chemical you still have to take into consideration what countries that chemical will be in, but one label works in them all (this is similar to the air pollution laws we are all familiar with, and how gasoline engines in machines like lawn mowers state if they are compliant with California regulations, even if they are sold in Arkansas). So the safety information associated with a chemical is influenced by the laws of the country which regulate the labeling of that chemical, and you should be prepared to look further.

PubChem LCSS

The National Institute of Health's (NIH) National Library of Medicine's (NLM) PubChem have developed LCSS that model the LCSS of the NRC, but extract data from multiple chemical compound databases. As of January 2021 there are LCSS for 150,8633 chemical substances that can be obtained through PubChem (not, the actual number will change over time and if you click that link you can both access the LCSS and see how many there are). This is a very valuable resource for finding safety information on chemicals and it is structured so you can understand where the information comes from. Why not take a minute to scroll through the LCSS of potassium permanganate, and there really is a lot of info there. Now to get a bearing on how complicated safety information can be, look at section 7 (First Aid), and note the first statement says Induce vomiting, but then in section 7.4 it says do not induce vomiting. Can you figure that out? Each of those statement provides links to the original source, and one was a solid that was inhaled, while the other was a liquid that was drunk.

COVID-19 Pandemic Issues

This course is being taught online because of the COVID-19 pandemic and there is a lot of conflicting information being spread across the internet, which makes it an ideal topic for discussing safety. In fact PPE that are normally used in the lab like masks and face shields are now often required in public. Students are expected to follow CDC guidelines when in public, which includes covering your mouth and nose when in public and it is worth watching this video from Florida Atlantic University and then discussing the kinds of PPE one should use to prevent either contact or air born spread of COVID-19.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 1.3.2: https://youtu.be/MMwYsGews-8

In the following video a chemistry professor at UC-Boulder who studies aerosol transmission of COVID-19 discusses the difference between water droplets and aerosols, while the video goes over some of the issues with evaluating online information. Later this semester we will learn more about aerosols, which are an interesting type of mixture called a colloidal dispersion. Aerosols are sort of a like mixture between a liquid and a gas, and the liquid components can absorb many chemical compounds but like a gas, travel great distances. Many so called "gas phase reactions" actually occur in aerosols and this are very important in fields like atmospheric chemistry.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 1.3.2: https://youtu.be/Y3zKRHzccMk

On Sep 21, 2020 the above video describes how the CDC posted an alert that the primary mechanism for the spread of COVID-19 was through aerosols and than strangly took that recommendation down saying it was posted in error, which can lead one to wonder what was going on with the CDC? Fast forward to Jan. 19, 2021 and the screen capture below clearly states that aerosols are involved in the spread of COVID-19. The important thing with regards to safety is to get all the information you can, assess where that information is coming from, and then make up your mind. Safety literacy is more than finding information on the web, it is also evaluating that information, and in the case of COVID-19, that could be the difference between life and death.

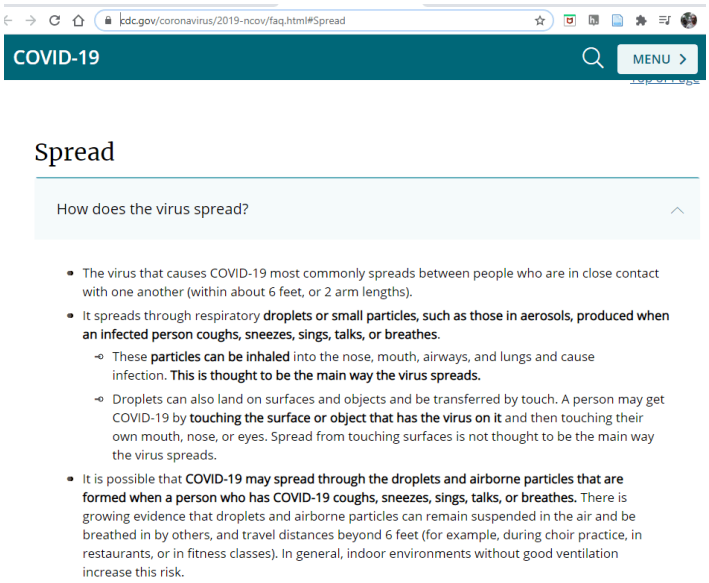


Figure 1.3.1: Screen capture CDC website, https://www.cdc.gov/coronavirus/2019-ncov/faq.html#Spread (01/19/2021)

Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Elena Lisitsyna

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1.4: Using ADAPT

Activity Overview

The goal of this activity is to learn how to upload PDF files of your work to the ADAPT platform of LibreText so it can be graded. You will make this two ways, by taking pictures with your cell phone and converting those to PDF, or by writing in a Google Doc and converting that to a PDF. There will be several steps, and note, if you have a scanner that converts to PDF you can use that instead of the Genius Cell phone app, but the images must be in PDF format. Each answer must be a different page of the PDF, and if they are hand written scans, you must print your name and date and then sign it. There are several steps to todays activity.

1. Write on a sheet of paper the following equation, and under it write today's date, print and sign your signature.

$$\%Salt = \left(\frac{m_S}{m_S + m_W}\right) 100 \tag{1.4.1}$$

2. Write on a new sheet of paper the following

I understand I need to be able to upload files to ADAPT and my grade will be influenced by the clarity of the files I upload. I will sign and date all scanned files to indicate they are my work. Under that statement write: *today's date, print and sign your name.*

3. Download and Install Genius Scan App to your cell phone (you may omit this if you have a scanner that will convert your images to a PDF file)
4. Take PDF scans of the your work in steps 1 & 2, with each problem being on a separate scan. If using Genius scan, feel free to make multiple images of each equation, and when you upload you can choose the best one to submit as your answer.
5. Go to ADAPT,
    - Upload the files as stated in the next section. Be sure to upload PDFs and not image files.

Upload Files to ADAPT

1. Log in to ADAPT

2. Select the course by clicking on its name

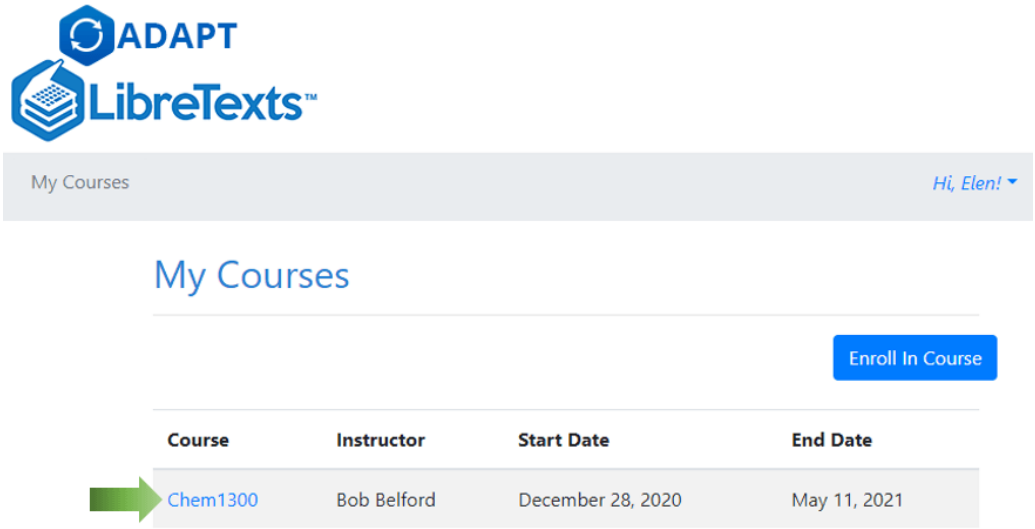


Figure 1.4.3: Select Chem1300

3. Select the Assignment you need to complete

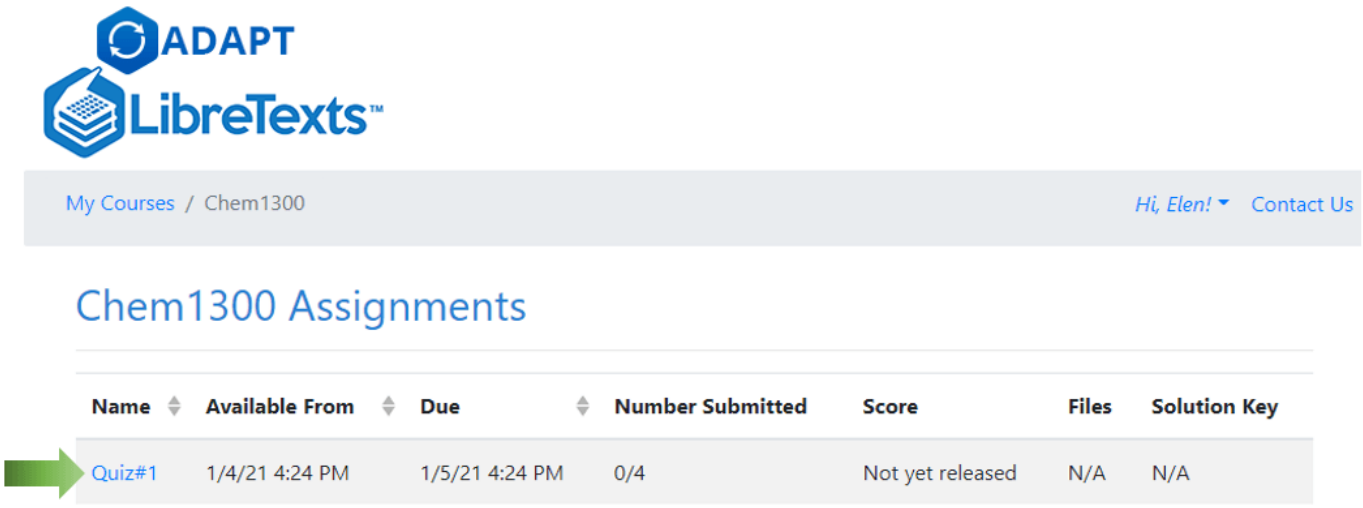


Figure 1.4.4: Select the assignment

4. Click "Upload New File". Depending on the type of a file you are uploading proceed to Upload as PDF or Upload as Image for further instructions

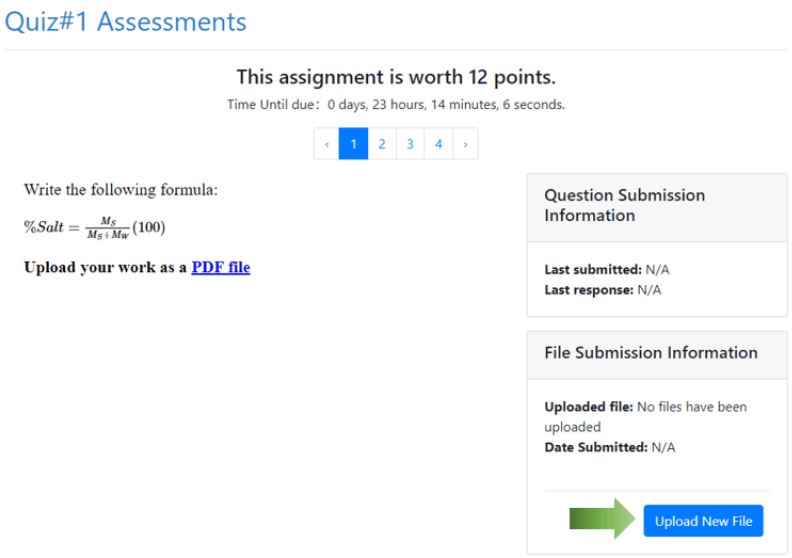


Figure 1.4.5: Upload New File

Upload as PDF

If you already have an app that can scan multiple pages as one document, and you are comfortable using it you can use it. If not, see section Genius Scan.

5. Select the "Upload your question file submissions from a full PDF that Adapt will cut up for you" option

6. Click "Browse" and select the PDF file you would like to upload

7. Click "Submit"

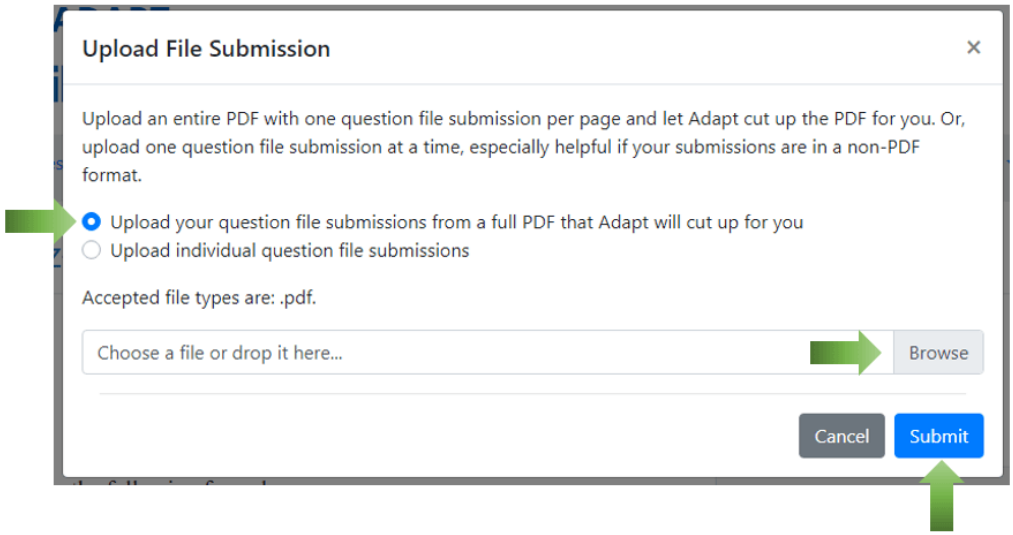


Figure 1.4.6: Upload PDF file

8. You can preview all pages of your PDF File by clicking on numbers in the PDF preview section. Type page number into the "Chosen cutups" box

9. Click "Set As Question File Submission". The page you've chosen will be displayed to a grader as your solution

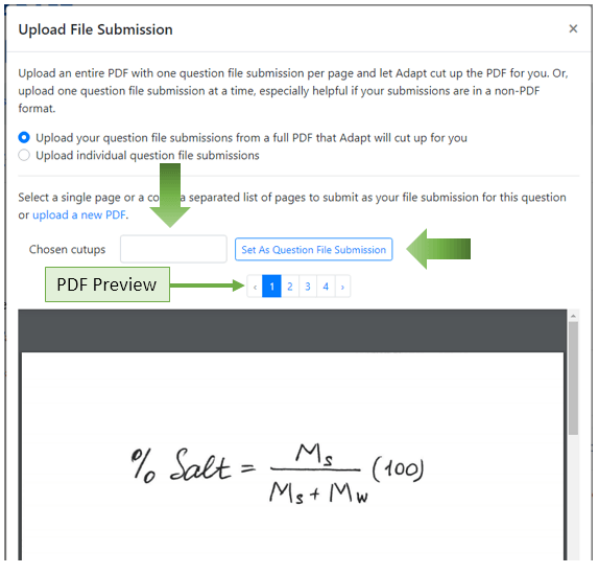


Figure 1.4.7: Choose cutup(s) for your file submission

10. Move to the next question and Click on Upload New File again. You don't have to upload your PDF file again.

11. Repeat steps 8-10

Upload as Image

5. **Take one picture per question.** If you have 4 questions in your quiz you need four pictures, each of them clearly showing your answer to one of the questions.

6. Select the "Upload individual question file submissions" option

7. Click "Browse" and select the image you would like to upload

8. Click "Submit". The image will be displayed to a grader as your solution

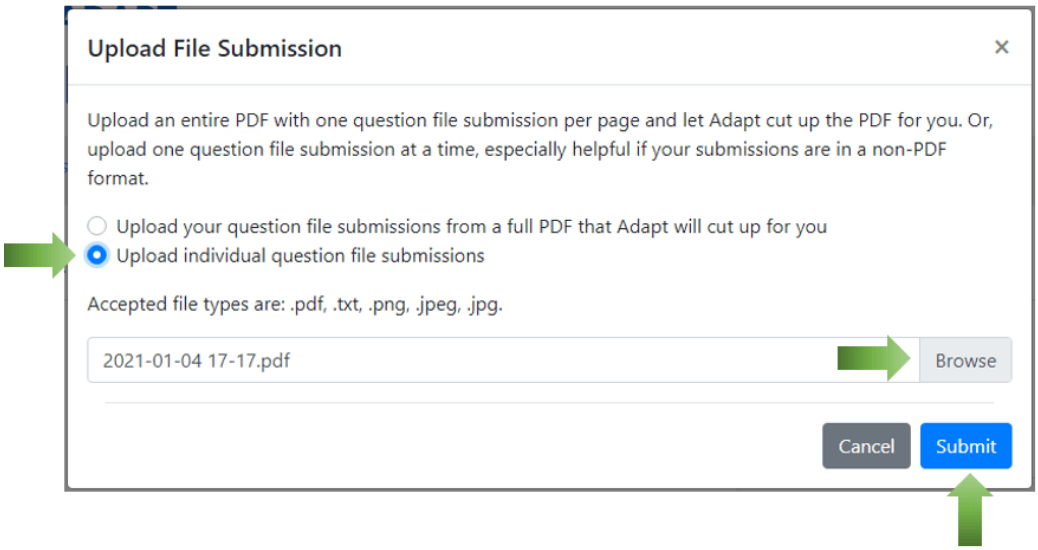


Figure 1.4.8: Upload an Image as your file submission

9. Move to the next question and Click on Upload New File again. You don't have to upload your PDF file again

10. Repeat steps 6-9 for each question

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2: MATHEMATICAL FUNDAMENTALS

In the introduction we pointed out that chemistry is an experimental science involving experiments that utilize both qualitative and quantitative data. The quantitative aspect of science means that scientific laws need to be understood in mathematically derived forms, and that you need to be able to solve algebraic problems and produce arithmetic solutions.



- Topic hierarchy

2.1: Mathematical Fundamentals

2.2: Algebra Review

2.3: Numbers

2.4: Units of Measurement

2.5: Measured Numbers and Significant Digits

2.6: Scientific Notation

2.7: SI Prefixes

2.8: Unit Conversions and Dimensional Analysis

2.9: Fractions and Percent

2.10: Graphing and Temperature Conversions

2.10.1: Graph Paper Page

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2.1: Mathematical Fundamentals

Commutative Property of Addition

Changing the order of addends does not change their sum

$$A + B = B + A \tag{2.1.1}$$
$$3 + 2 = 5 \;\; while, \;\; 2 + 3 = 5$$

note, subtraction is not commutative

$$A - B \neq B - A \tag{2.1.2}$$
$$3 - 2 = 1 \;\; while, \;\; 2 - 3 = -1$$

Commutative Property of Multiplication

$$AB = BA \tag{2.1.3}$$

Note division not commutative

$$\frac{A}{B} \neq \frac{B}{A} \tag{2.1.4}$$
$$\frac{3}{2} = 1.5, \;\; while \;\; \frac{2}{3} = 0.67$$

Associative Property of Addition.

$$(A + B) + C = A + (B + C) \tag{2.1.5}$$
$$\text{left side: } \; (2 + 3) + 4 = 5 + 4 = 9$$
$$and$$
$$\text{right side: } \; 2 + (3 + 4) = 2 + 7 = 9$$

Associative Property of Multiplication.

$$2(3x4) = (2x3)4 \tag{2.1.6}$$
$$\text{left side: } \; 2(12) = 24$$
$$and$$
$$\text{right side: } \; (6)4 = 24$$

Distributive Property of multiplication .

$$A(B + C) = AB + AC \tag{2.1.7}$$
$$\text{left side: } \; 2(3 + 4) = 2(7) = 14$$
$$\text{right side: } \; 2(3) + 2(4) = 6 + 8 = 14$$

note the reverse process allows you to factor out a value if you want to solve for it, so in the next equation, say you want to solve for T

$$m_1 c_1 T + m_2 C_2 T = Q \tag{2.1.8}$$
$$T(m_1 c_1 + m_2 c_2) = Q$$

and you can now solve for T

$$T = \frac{Q}{m_1 c_1 + m_2 c_2} \tag{2.1.9}$$

Exponentiation

Exponentiation is the repetition of multiplication

$$2^n \text{ is 2 times 2 repeated n times} \qquad (2.1.10)$$
$$2^4 = 2(2)(2)(2) = 16$$

Roots of a number

The root of a number to the power of n is the number that when multiplied n times equals the original number. The root of a number is actually taking the number to the power of the reciprocal of the exponent: the third root of 64 is 64 to the power of the reciprocal of three.

$$\sqrt[3]{64} = 64^{\frac{1}{3}} \text{ means what number when multiplied by itself 3 times equals 64} \qquad (2.1.11)$$
$$\sqrt[3]{64} = 4 \text{ because } 4(4)(4) = 64$$

Order of Operations: PEMDAS

PEMDAS is an acronym standing for **P**arenthesis **E**xponents **M**ultiplication **D**ivision **A**ddition **S**ubtraction, and used to remember the order of performing operations for calculations involving multiple arithmetic operations. The PEMDAS convention that is used in both math and computer science is to perform operations in the following order

- P first compete all calculations in **parenthesis**
- E second complete all calculations involving **exponents** (and roots - which are essentially exponents)
- MD third perform all calculations involving **multiplication** and **division**, typically going left to write (these are of priority, just as roots and exponents are)
- AS last addition and subtraction, typically going left to write (these are of equivalent priority)

Consider the following operation

$$2(3+4) \qquad (2.1.12)$$

If you incorrectly performed operations left to right you would multiple 2 time 3 and come up with 6 and then add 4 giving the incorrect value of 10.

the proper sequence is to add the 3+4 to get 7 and then multiply by two giving the correct solution of 14.

That is

$$2(3+4) \neq 2(3)+4$$

$$2(7) \neq 6+4$$

$$14 \neq 10$$

The correct answer for $2(3+4)$ is 14.

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2.2: Algebra Review

Introduction

In section 1.2 (Mathematics and Scientific Communication) we went over how algebraic equations can be derived from theoretical equations, and these algebraic equations can be solved for a an unknown variable. Once this is done, the mathematical values of the known variables can be substituted in and the equation can be arithmetically solved for a numerical answer.

We are now going to look at some mathematical operations that can be performed on both algebraic and arithmetic equations that maintain an equality.

Lets look at a simple equivalence relationship. If A= 6 calories and B= 6 calories, then algebriacally,

A=B

Algebraic Operations that Maintain an Equivalence

The following operations are commonly used to solve algebraic equations for specific variables. For these to work, they must maintain the equivalence statement.

1. Adding the same value to both sides of an equivalence statement

if

$$A = B$$

then

$$A + C = B + C \tag{2.2.1}$$

2. Subtracting the same value to both sides of an equivalence statement

if

$$A = B$$

then

$$A - C = B - C \tag{2.2.2}$$

3. Multiplying same value to both sides of an equivalence statement

if

$$A = B$$

then

$$A(C) = B(C) \tag{2.2.3}$$

(we often use parenthesis to indicate multiplication, but sometimes omit them)

AC=BC

4. Dividing same value into both sides of an equivalence statement

if

$$A = B$$

then

$$\frac{A}{C} = \frac{B}{C} \tag{2.2.4}$$

5. Multiplying and dividing by the same number does not change the value.

i

$$\begin{gathered} C = A(B) \\ then \\ C = A(B)(\frac{D}{D}) \\ so \\ C = AD\left(\frac{B}{D}\right) = BD\left(\frac{A}{D}\right) = AB \end{gathered} \tag{2.2.5}$$

Test Yourself

Homework: Section 2.2

Activity 2.2.1

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this Interactive Content online

NOTE: The step by step worked out solutions to these questions can be found in homework section 2.2 Algebra Review

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- Vincent Belford (H5P interactive modules)

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2.3: Numbers

Introduction

Scientists use three basic types of numbers when quantifying matter; counted numbers, measured numbers and defined numbers.

- **Counted Numbers:** A counted number is an exact number in the sense that you have an entity, and you count the number of entities. So an exact number has two parts, the entity, and the number of entities. If you have 16 rocks in your hand, you have sixteen rocks, not more, not less. If you break one of the rocks into two, you now have 17 rocks in your hand, not more, not less. These are exact numbers. **Counted numbers are exact and there is no uncertainty about their value.**

$$\underbrace{16}_{number}\ \underbrace{rocks}_{entity}$$

Counted Numbers are Exact

- **Measured Numbers:** Now the mass of the rocks is a measured number and requires a third aspect to describe, which is the unit, defined by the scale used to measure it. So a measured number has 3 parts; magnitude, unit and entity. Sixteen one pound rocks weighs less than one 500 pound rock, although the number 16 is more than the number one, and this is because different units are being used to describe the mass. If you break each of the sixteen one pound rocks in half, you have 32 rocks, but they still weigh the same. **Measured numbers have uncertainty that is indicated by their number of significant digits** (section 2.5).

$$\underbrace{16}_{number}\ \underbrace{grams}_{unit}\ \underbrace{rock}_{entity}$$

Measured Numbers have Uncertainty

- **Defined Numbers:** A defined number has a value inherent in its definition, and like a counted number is an exact number, but it may, or may not have uncertainty. Defined numbers are often used in measurements, for example, twelve inches is defined to be one foot and there is no uncertainty in that equivalence.

$$12\text{inches} \equiv 1\text{foot}$$

- But not all defined numbers are integers, like the number π, which is defined to be ratio of the circumference to the diameter of a circle.

$$\pi \equiv \frac{circumference}{diameter}$$

  π is an exact number in the sense that it is exactly the ratio of the circumference to the diameter of a circle, but it is an irrational number and like a measured number, when we write down the value of π we give it a precision (3.14, 3.142, 3.1416, 3.14159,...), which defines how precisely we express the number.

Understanding defined numbers is of importance, and as we shall learn, as of May 20, 2019, the basic units of the metric system became based on defined numbers (2019 SI unit definitions).

Expressing Numbers

There are four common ways to express numbers, and some of these are preferable when the numbers are very large or very small.

1. Floating Decimal Notation Number
2. Scientific Notation
3. SI Prefixes
4. Logarithms

Floating Decimal Numbers

Floating numbers are common numbers where the "magnitude" is determined by the position of the decimal point. Digits to the left of the decimal are larger than 1 and digits to the right are less than one, or a fraction. So 12.121200 and 121212.00 are two numbers of identical digits but the second is 10^4 times larger than the first . A fraction typically has a zero in front of the decimal, so 0.25 is $\frac{25}{100}$ or one quarter $\frac{1}{4}$.

Scientific Notation

Unlike a floating numbers scientific notation has two parts, first a number with only one digit to the left of the decimal and a variable number of digits to the right of the decimal, which is multiplied by 10 to a power, such that it equals the value of the desired number. We call the power of 10 its exponent.

'_' allowed only in math mode

Where n is the power and the underscore represents the number of digits in the number. Note, due to this format, scientific notation always has an unambiguous number of significant digits.

Q&A: What is Exponentiation?

Exponentiation is the replication of multiplication the way multiplication is the replication of addition

$$10^0 = 1$$

$$10^1 = 10 \qquad 10^{-1} = \frac{1}{10} = 0.1$$

$$10^2 = 10(10) = 100 \qquad 10^{-2} = \frac{1}{10^2} = \frac{1}{100} = 0.01$$

$$10^3 = 10(10)(10) = 1000 \qquad 10^{-3} = \frac{1}{10^3} = \frac{1}{1000} = 0.001$$

Q&A: Can the exponent have units?

No, the exponent can not have units, it is the number of times you multiplied or divided the number by 10, but you write the units behind the number expressed in scientific notation.

$$1.333 \times 10^4 \ L = 13,330 \ L \tag{2.3.1}$$

Advantages of Scientific Notation

1. Allows Awkwardly Large and Small Numbers to Be Expressed in Term of Compact and Easily Written Numbers
2. Allows Accurate Representation of the Number of Significant Figures in a Number, That Is a Measurement's Precision, the "Certainty" of Our Measurements

SI Prefixes

SI prefixes are similar to scientific notation, where it is common to express a number with between 1 and 3 digits to the left of the decimal point, times 10 to the power of 3, but use the appropriate SI prefix to represent 10 to the power of 3.

$$_.__ \times 10^{3n} \tag{2.3.2}$$

$$__.__ \times 10^{3n} \tag{2.3.3}$$

$$___.__ \times 10^{3n} \tag{2.3.4}$$

Table 2.3.1: Some common SI prefixes.
(n=1,2,3,-1,-2,-3)

10^{3n}	Name	Abbreviation		Name	10^{-3n}
10^{+3}	kilo	k	m	milli	10^{-3}
10^{+6}	mega	M	μ	micro	10^{-6}
10^{+9}	giga	G	n	nano	10^{-9}

Note, it is common to express floating numbers with a comma every three decimal places (power of 10^3). So the numbers

$$1,330,000,000m = 1.33 \times 10^9 m = 1.33\ gigameters \tag{2.3.5}$$
$$13,300,000,000m = 13.3 \times 10^9 m = 13.3\ gigameters \tag{2.3.6}$$
$$133,000,000,000m = 133 \times 10^9 m = \ 133\ gigameters \tag{2.3.7}$$

Note

Not all SI prefixes are in magnitudes of 10^3, the complete list is in section 2.7 and you will be responsible for memorizing the ones in red. This includes common ones between a thousand and a thousandth, like centi, deci... You should make flash cards of the prefixes in section 2.7 now, so you do not have to memorize them when we cover them.

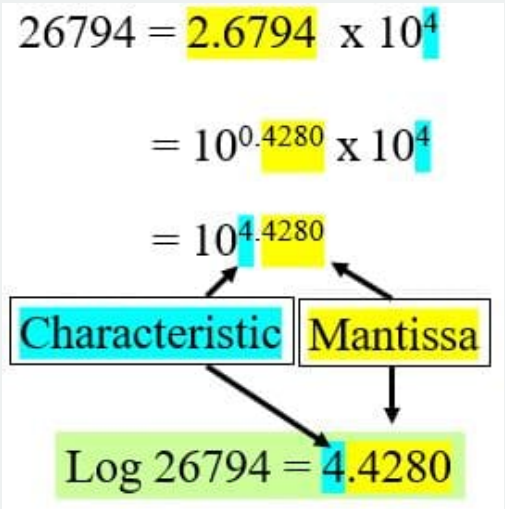
Logarithms

Logarithms are very similar to scientific notation, but the pre-exponential is expressed as 10 to the power of a fraction, and the log of a number, is the power of 10 that equals a number. Logarithms are often not covered in general chemistry 1, but they are extensively used in general chemistry 2.

Deeper Look: Logarithms

This is optional material for general chemistry 1, but you will need to understand this for general chemistry 2.

Let's consider the number 26794. It can be expressed in scientific notation or as a log



So logarithms are very similar to scientific notation, where the number being multiplied by 10 to a power is expressed as 10 to the power of a fraction. Please see the math review in Chapter 14 prelude for

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2.4: Units of Measurement

Metrology

Metrology is the science of measurement. There are two fundamental tenants to metrology; traceability and uncertainty. Traceability is the ability to relate a measurement to a standard and uncertainty deals with how certain you are in the measurement based on the practice and technique you use to make the measurement. A thorough coverage of units and metrology is given in the Chem 1402 General chemistry 1 section on Units of Measurement.

Lets start with something we are all familiar with, like the units of length (figure 2.4.1). Some are very small compared to others and we would choose a scale that is relevant to the magnitude of the entity we are measuring, like a map would relate distance from Little Rock to Memphis miles and not inches, but you would probably not want to know the length of your hand in miles.

Ångstrom	1E−6 cm
Astronomical unit	1.495E8 km
Bolt (U.S., cloth)	120 ft
Cable length	720 ft
Chain (engineer's)	100 ft
Chain (surveyor's)	0.1 furlong
Cubit	18 in
Ell	45 in
Foot (U.S.)	12 in
Foot (British)	0.4 pace
Furlong	1/8 mile
Hand	4 in
Inch (U.S.)	1/12 ft
Inch (British)	1/36 yard
League	3 miles
Light year	9.46E12 km
Mile (U.S., staute)	5280 ft
Mile (nautical)	1.853 km
Nail (British)	2.25 in
Pace (British)	30 in
Parsec	3.084E13 km
Point (printer's)	1/72 in
Rope (British)	20 ft
Span	9 in

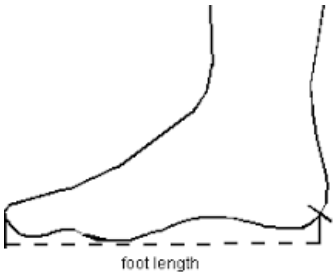


Figure 2.4.1: On the left are some of the units to describe length, some of which were based on human body parts (foot) or actions (pace).

The earliest units were based on artifacts, some of which were human body parts (foot) or action (pace). So if a king decreed that his foot was the standard of length, then someone could make a standard by marking an object by the length of the king's foot and use that object to measure other objects. But what happens if the king dies and the new king's foot is longer or shorter than the old kings? In fact, it is worst than that, as the king's foot can change in length, depending on how old he is, how he stands, etc.

Today science uses what we commonly call the metric system. In fact most countries in the world use the metric system as the normal units of measurement.

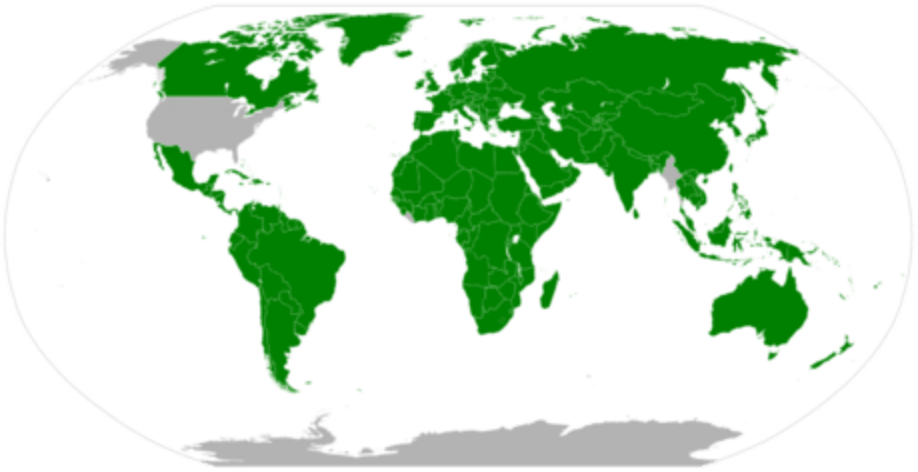
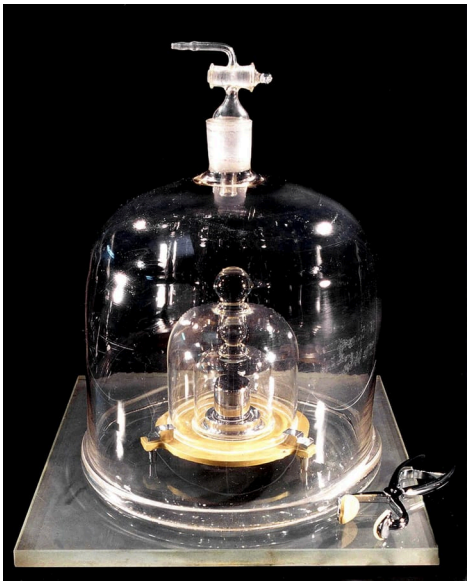
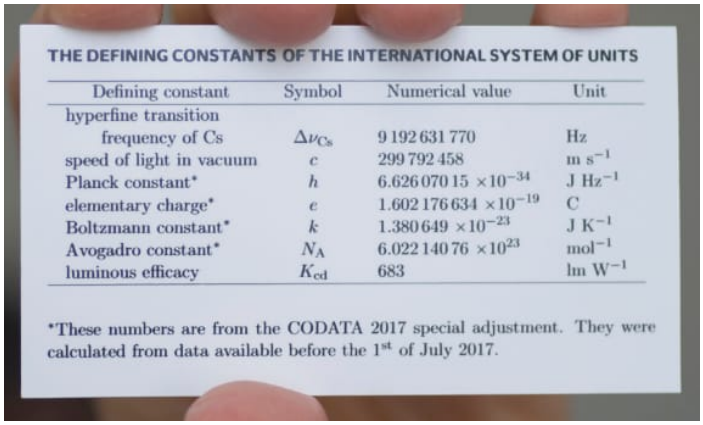


Figure 2.4.2: Areas of the world using the metric system.

The customary units of measurement in the US are the foot, inch and pound, even though most of the rest of the world uses the metric system, as does the scientific community. The metric system has evolved over the years and today science uses the ***Système international d'unités* (SI)** units of measurement, which are far more complicated than the original metric units, and there are 7 fundamental SI base units that all matter can be represented under. To give a feel for the magnitude of the challenges involved the base units of the SI system underwent a change on May 20, 2019, and are now based on defined numbers, the 7 fundamental constants of nature. The details of these definitions are above the level of this class, in fact they are above the level of the typical general chemistry class, but if you are interested you may want to peruse through the section 1B.1.3.1.2 of the gen chem 1 LibreText.





THE DEFINING CONSTANTS OF THE INTERNATIONAL SYSTEM OF UNITS

Defining constant	Symbol	Numerical value	Unit
hyperfine transition frequency of Cs	$\Delta\nu_{Cs}$	9 192 631 770	Hz
speed of light in vacuum	c	299 792 458	m s^{-1}
Planck constant*	h	$6.626\,070\,15 \times 10^{-34}$	J Hz^{-1}
elementary charge*	e	$1.602\,176\,634 \times 10^{-19}$	C
Boltzmann constant*	k	$1.380\,649 \times 10^{-23}$	J K^{-1}
Avogadro constant*	N_A	$6.022\,140\,76 \times 10^{23}$	mol^{-1}
luminous efficacy	K_{cd}	683	lm W^{-1}

*These numbers are from the CODATA 2017 special adjustment. They were calculated from data available before the 1st of July 2017.

Figure 2.4.2: On the left is an image of the old standard kilogram (credit: BIPM), and on the right are the 7 defined constants that are considered as "invariants of nature" and thus are numbers (credit: Stoughton/NIST). How these are implemented are above the level of this course, and what is important is that students realize that as science advances, the most fundamental concepts as the values of the units of measurement evolve.

SI Base Units

The seven SI base units are listed in Table 2.4.1. These can be combined to form derived units that describe any measurable physical quantity. Thus by defining these seven SI base units, any physical measurement can be reproduced. In this class we will use the first five units extensively; the meter, kilogram, second, kelvin and mole (and general chemistry 2 will also use the ampere).

Table 2.4.1: Seven SI Base Units. Chem 1300 students need to know the units in red.

length	**meter**	**m**
mass	**kilogram**	**kg**
time	**second**	**s**
temperature (absolute)	**kelvin**	**K**
amount of substance	**mole**	**mol**
electric current	ampere	A
luminous intensity	candela	cd

Note, the SI unit for mass is the kilogram, represents 1,000 grams, and uses the si prefix kilo, which we will cover in section 2.7 (SI Prefixes).

Derived SI Units

As state above, any measurable phenomena can be described in terms of the SI base units. There are two types of described units; those that are combinations of base units, like $m^2$for area, and those with their own unit symbols, like Joules (J) for energy.

$$\underbrace{1J = 1\frac{kg \cdot m^2}{s^2}}_{\text{SI Derived Unit of Energy in terms of SI Base Units}} \tag{2.4.1}$$

Table 2.4.2 shows common derived SI units:

Table 2.4.2: Derivied SI units and their dependencies on the SI bas units.

Quantity	Unit	Explanation
Work	N	Newton = kg m s^{-2}
Pressure	Pa	Pascal = N m^{-2}
Energy	J	Joule = N.m
Charge	C	Coulomb = A.s
Electric Potential	V	Volt = J/C
Power	Watt	1 watt = 1 J/s
Volume	L	1L = $1dm^3$
Volume	mL	1mL = $1cm^3$
Molarity	M	mol/L

Note, there are many non-SI units like the calorie, atm, foot, etc., which can also be described in terms of the SI base units. Conversions between units is covered in section 1B.3.5.

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2.5: Measured Numbers and Significant Digits

Uncertainty

Measured numbers have "uncertainty" in that one can not know the value of a measurement that is smaller than the smallest unit of the scale. Precision is a measure of how close successive measurements are to each other. Precision is influenced by the scale, and when reporting a measurement, you report all certain values, and the the first uncertain one (which you "guesstimate"). This is illustrated in Figure 2.5.1. The scale on the left is a cm scale because the smallest value you know is in cm, and marker (arrow) is clearly greater than 1 and less than 2 centimeters, and so would be reported as 1.6cm, or maybe 1.7cm (as you report all certain values, plus the first uncertain value). The scale on the right is a mm scale, and you know the marker is greater than 16 mm and less than 17mm, and you would report it as 1.67cm (which is the same as 16.7mm).



Figure 2.5.1: A cm scale (left) and mm scale (right).

So if 100 people measure the same object, they will come up with different values, and the closeness of those values is dictated by the scale they use. The mm scale is more precise because everyone would come up with values between 1.6 and 1.7 cm, while with the cm scale, their values would be between 1 and 2 cm.Significant Digits

Significant Digits, Exact Numbers and Defined Numbers:

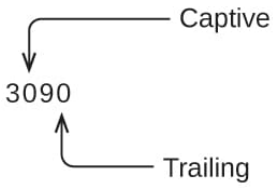
A counted number is an integer and thus is an exact number, for which there is no uncertainty. So it does not influence significant figures. If one looks on the web one often sees people saying that a counted number has an infinite number of significant figures, mathematically that may work, but it is incorrect, in that you do not need an infinite number of significant digits to exactly define a counted number. The fact is, there is no uncertainty in an exact number, 3 cows is 3 cows. Now defined number may or may not have significant digits. Twelve inches = 1 foot, there are no significant digits. But an irrational number like π, which is also a defined number, would require an infinite number of significant digits to exactly define, and if you use pi in a calculation, you should use enough significant digits so that it does not determine the number of significant digits in your final answer. Simply speaking, significant digits are a way to indicate the precision of measured values, and the above rules enable you to preserve them in calculations. Exact and defined numbers do not involve measurements, and so they do not influence the number of significant digits in a calculations (unless you do not use enough digits for a defined value like π).

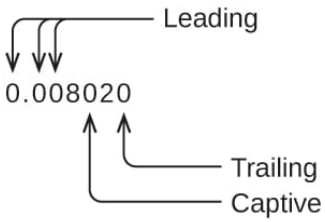
Significant Digits

It should be noted that we use the terms "significant digits" and "significant figures" synonymously. In making measurements we report all certain values and the first uncertain value and the number of digits in that number is the number of significant figures in the answer. First we will go over the rules for representing significant digits in a reported measured value and then the rules for determining them from a measurement. Then we will go over how to carry significant digits in a calculation and how to round up or down (truncate) calculated answers that have more digits than are significant.

Significant Digits Rules

Whenever you make a measurement properly, all the digits in the result are significant. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Well, for starters, all nonzero digits are significant, and it is only zeros that require some thought. We will use the terms “leading,” “trailing,” and “captive” for the zeros and will consider how to deal with them.

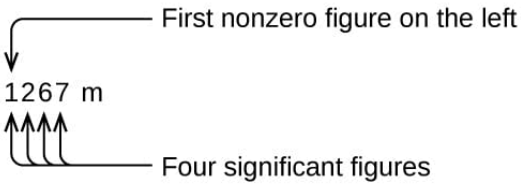


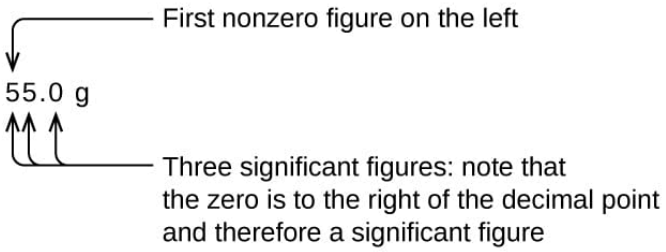


1. Non Zeros are always significant
2. Leading Zeros are never significant
3. Captive Zeros are always significant
4. Trailing Zeros are only significant if a number has a decimal point*

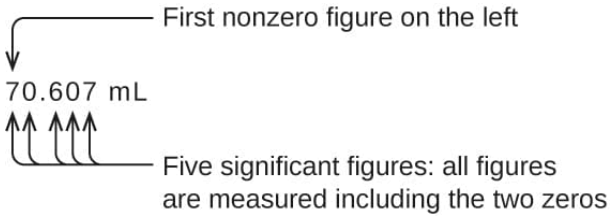
*Note: Some textbooks/instructors will state the number of significant digits is undetermined if you have trailing digits without a decimal point, so 200 could have one, two or three sig figs. In my class we will say 200 has one significant digit, and you can't express it two or three sig figs in floating decimal point notation, and that is why we sometimes need to use scientific notation.

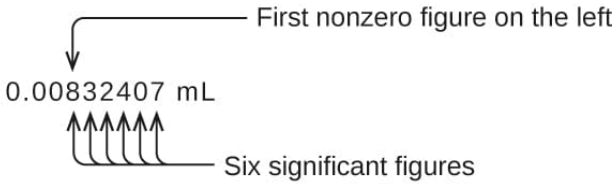
Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point.





Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.





The leading zeros in this example are not significant. We could use exponential notation (as described in Appendix B) and express the number as 8.32407×10^{-3}; then the number 8.32407 contains all of the significant figures, and 10^{-3} locates the decimal point.

Note

Remember, Significant Digits apply to measured numbers, not counted or exact numbers. 100 mushrooms does not have any significant digits, it is an exact number, but 100.0 g of mushrooms has 4 significant digits, as they were measured with a scale where 1 gram was the smallest increment. Some instructors say the have an infinite number of significant digits, which implies that they do not contribute to the number of significant digits in a calculation.

Reporting Significant Digits for Measured Numbers

Significant Digits (Figures) are a way of expressing the uncertainty in a reported value, where you indicate all certain and the first uncertain digit. The goal is to:

- Report all certain values
- Report first uncertain value
  - Uncertain Value is a "guess" between smallest unit of scale
  - Successive Measurements will vary by uncertain value

Classroom Activity

Go to the following Google Forms and Fill them out. Please do not do this if you are not in my class. You need to write down your answer on a sheet of paper, as we will use it in a Google Doc.

Decimeter Scale

Centimeter Scale

Millimeter Scale

Exercises 2.5.2 through 2.5.5 are measurements of the same piece of paper using different scales.

Exercise 2.5.1

What is the height and width of the following piece of paper in units of meters?

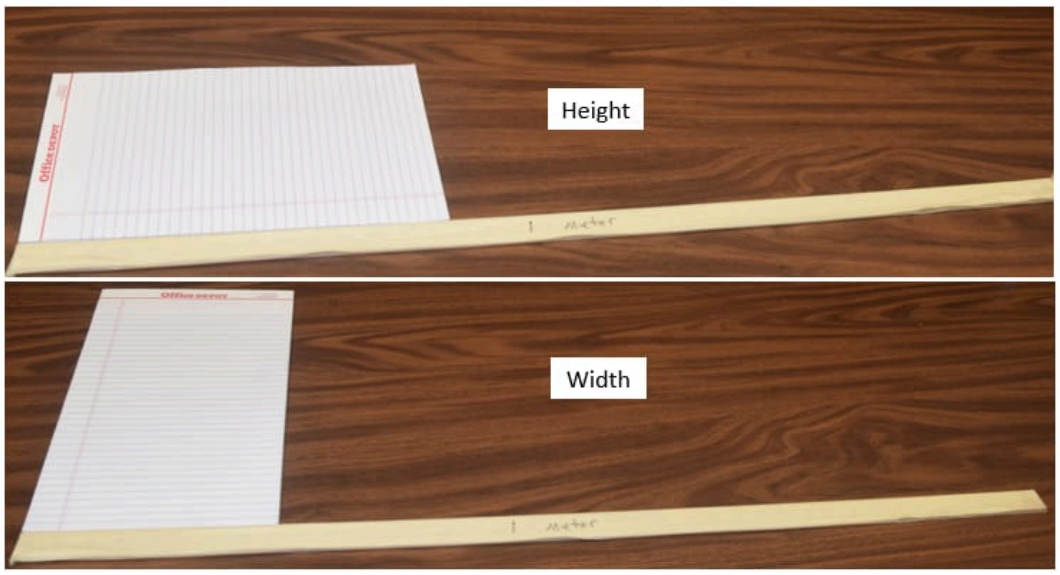


Figure 2.5.2: Sheet of paper measured with a meter scale, see section on SI Units. (Copyright; Bob Belford, LibreText)

Answer available at the end of this section.

Exercise 2.5.2

What is the height and width of the following piece of paper in units of centimeters?

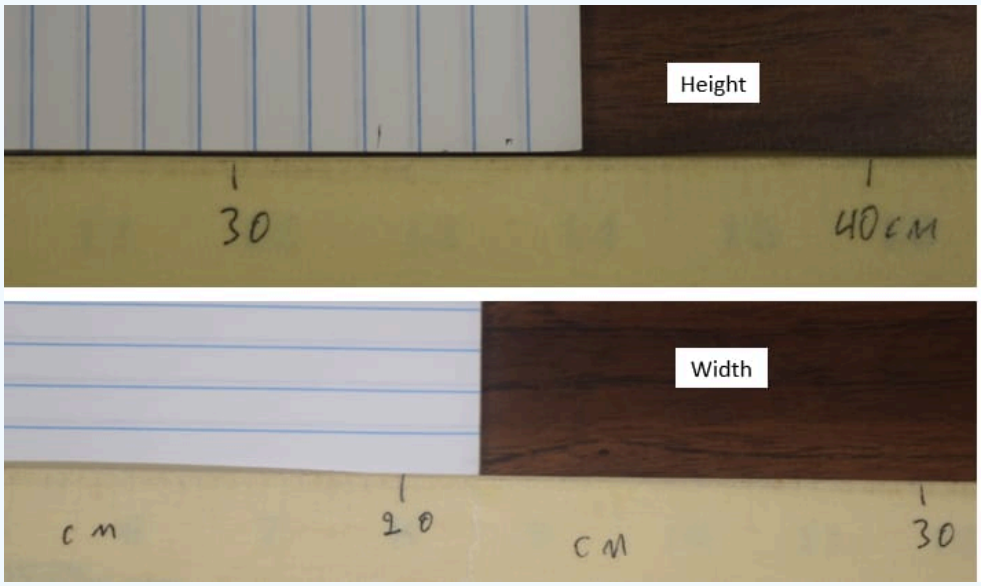


Figure 2.5.3: Sheet of paper measured with a centimeter scale (see section on SI Units. (Copyright; Bob Belford, LibreText)

Answer available at the end of this section.

Exercise 2.5.3

What is the height and width of the following piece of paper in units of centimeters?

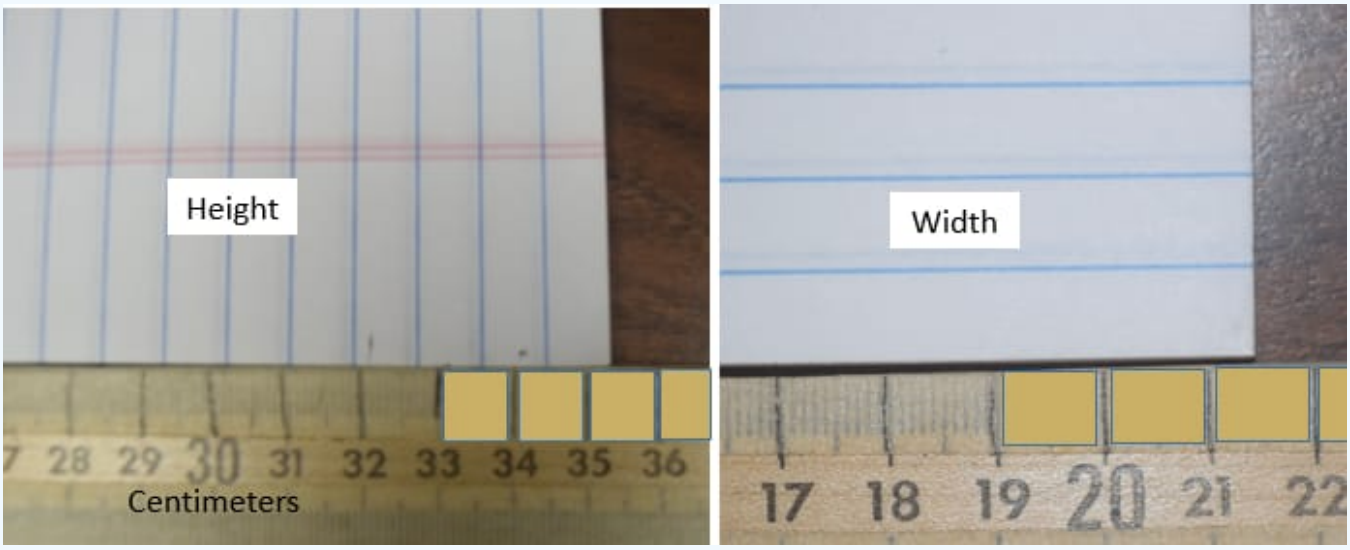


Figure 2.5.4: Sheet of paper measured with a centimeter scale (see section on SI Units. (Copyright; Bob Belford, LibreText)

Answer available at the end of this section.

Exercise 2.5.4

What is the height and width of the following piece of paper in units of centimeters?

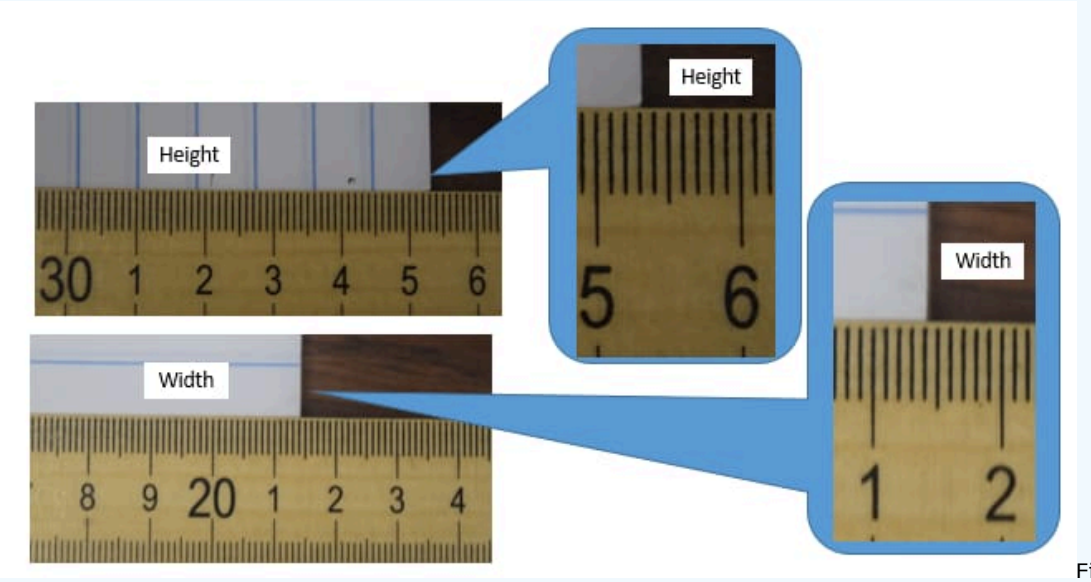


Figure 2.5.5: Sheet of paper measured with a centimeter scale (see section on SI Units. (Copyright; Bob Belford, LibreText)

Answer available at the end of this section.

Significant Digits, Calculations and Rounding off Answers

It is important to be able to preserve the number of significant Figures in calculations based on the precision of the data used in those calculations, and there are two sets of rules, one for multiplication and division, and the other for addition and subtraction.

Multiplication and Division

The number of significant Figures in the product of multiplication or dividend of division is based on the valued used with the least number of significant digits.

Exercise 2.5.5

How many significant Figures are in the answer to the following problem?

$$\frac{(3.42)(127)}{1.6} = ?$$

Answer available at the end of this section.

Why are Leading Zeros Not Significant?

Consider you weigh 1000 paper clips, accurately counted and the combined weight is 1 gram. This has one significant digit. If you then asked how much did each paperclip weight, you divide 1g by 1000 paper clips to get the mass of each paper clip

$$1000\ paper\ clips = 1g \tag{2.5.1}$$

$$\frac{1000\ paper\ clips}{1000\ paper\ clips} = \frac{1g}{1000 paper\ clips}$$

$$\frac{\cancel{1000\ paper\ clips}}{\cancel{1000\ paper\ clips}} = \frac{1g}{1000 paper\ clips}$$

$$1 = 0.001 \frac{g}{paper\ clip}$$

$$1\ paper\ clip = 0.001g$$

Since 1000 paper clips is an exact number and does not have significant digits, and 1 gram has one significant digits, the answer would have one significant digits, and thus the convention that leading zeros are never significant. if we said 1000 paper clips = 1.00 grams, the one paper clip = 0.00100 g.

Addition and Subtraction

Here, the number of significant Figures is based on the least precise measured value. In this case you actually have to line the numbers up by the decimal point and do the math, and then limit the answer to the precision of the least precise value.

Consider:

$$\begin{array}{l} 1.2 \\ \underline{4.41} \\ 5.61 \end{array}$$

↑ limit final answer to the tenths column: 5.6

Exercise 2.5.6

How many significant Figures are in the answer for 99 + 1?.

Answer available at the end of this section.

Note: in the above exercise, a number with two sig. figs was added to a number with one sig fig., and the answer had three sig. figs. Unlike multiplication and division, you must do the math to determine the number of sig. figs for addition and subtraction.

Rounding Off Answers

When you have to truncate an answer by removing numbers you may round up or round down. Truncation is removing a digit from the number, and all digits smaller than that digit.

- If truncated digit is less than 5, round down (use value of preceding digit)
- if truncated value is greater than 5, round up (add one to preceding value).
- if truncated value equals 5, look at the preceeding digit **(Even Down Odd Up)**.
  - if it is even or non existent (0,2,4,6,8), round down (treat the 5 as a 4)
  - if it is odd (1,3,5,7,9), round up (treat the 5 as a 6)

Warning!

If the truncated value is 5 some instructors and textbooks say to always truncate up (0,1,2,3,4 round down, while 5,6,7,8,9 round up) so you should **ask you instructor what they want you to do in their class, and do that!** The problem with always rounding 5 up is the accumulation of round off error when rounded up numbers are used in subsequent calculations like averaging. Five is half way between 0 and 10 and so statistically it would be better to round up half the time and round down the other. Since there are five even numbers (0,2,4,6,8) and five odd numbers (1,3,5,7,9), we round the evens down and the odds up. If five is the last given number, look at the number that preceeds it and round down if even and up if odd, so the result is always rounded to an even number.

Table 2.5.1: Examples of rounding-off

number to round / no. of sig. digits	result	comment
round 34.216 to 3 sig figs	34.2	First non-significant digit (1) is less than 5, so number is simply truncated.
round 2.252 to 2 sig figs	2.2	First non-significant digit is 5, since the preceding digit is even, round down (treat the 5 as a 4)
round 2.353 to 2 sig figs	2.4	First non-significant digit is 5, since the preceding digit is odd, round up (treat the 5 as a 6)
round 2.25 to 2 sig figs	2.2	First non-significant digit is 5, since the preceding digit is even round down
round 39.99 to 3 sig figs	40.0	Crossing "decimal boundary", so all numbers change.
round 85,381 to 3 sig figs	85,400	The two zeros are just placeholders
round 0.04597 to 3 sig figs	0.0460	The two leading zeros are not significant digits.

Exercise 2.5.7

What is the length in centimeters of 3 sheets of paper aligned in a row ig the width of each sheet was obtained from the measurement and scale used in exercise 2.5.3 (width = 21.3 cm). Assume the three sheets are exactly touching with no overlap or space between them.

Answer available at the end of this section.

Exercise 2.5.8

What is the height in centimeters of 3 sheets of paper aligned in a column if height of each sheet was obtained from the measurment and scale used in exercise 2.5.3 (height = 35.2cm cm). Assume the three sheets are exactly touching with no overlap or space between them.

Answer available at the end of this section.

Compound Calculations

Use all digits during intermittent steps and only truncate the final answer. You have to manually determine the number of significant digits of intermediate addition or subtraction steps (but use all digits in subsequent calculations). Do not truncate (round off) intermediate steps

Example 2.5.3

$$\frac{99.0g}{1g+99g} = ? \tag{2.5.2}$$

Solution

0.990g because the rule of addition and subtraction give three significant digits for the denominator, thus both the denominator and numerator have three significant digits.

Perform addition in the parenthesis first. **Note the following problem will give the wrong answer if you round off during intermediate calculations.**

Example 2.5.4

(1.89g + 1.55g + 1.0g) x 7.57g =

Solution

(1.89g + 1.55g + 1.0g)= 4.44
4.44 x 7.57=33.6108=34.

Note, if you truncated 4.44 to 4.4 you would have gotten the wrong answer!

(1.89g + 1.55g + 1.0g)= 4.44 = 4.4 (round down), if you then used that value for the subsequent step you would get the wrong answer

4.4x7.57=33.308=33 (wrong answer)

You only truncate in the last step.

Exercise 2.5.9

What is the area of the rectangle formed by 9 sheets of paper forming 3 rows and 3 columns, where the rows are the width of the sheet of paper in exercise 7 (21.3 cm) and the height of the columns are made by three sheets with the height of exercise 8 (35.2 cm). Assume the three sheets are exactly touching with no overlap or space between them.

Answer available at the end of this section.

Note

If in exercise 2.5.9 you had used the answers from exercises 2.5.7 and 2.5.8 you would have gotten the wrong answer as you rounded up during an intermediate calculations. **Only round up or down on the final step. This may mean you need to make a note of the number of significant digits of numbers used in intermediate steps.**

Contradictions of the Rules

Sometimes the rules break down and you need to be aware of it. For example, if you weight two nickels and each nickel weights 5.0 grams, what is the weight of both nickels? Well, multiplication is the repetition of addition and so you could multiply the weight of each nickel by two

$$2\,\text{nickels}\left(\frac{5.0\,\text{grams}}{\text{nickel}}\right) = 10.\ \text{grams} \tag{2.5.3}$$

or you could add the mass of the two nickels

$$5.0\,g + 5.0\,g = 10.0\,g \tag{2.5.4}$$

So when you multiplied the answer would have two significant Figures because the number with the least was the measured mass (the 2 nickels was an exact counted number), but when you added, they were all known to the tenths position, and so your answer has 3 significant digits.

Experimental Error

There are two basic types of error, systematic and random error. Systematic error is often due to faulty equipment or techniques. For example, if a meter stick was worn down on the end that reads zero, it would give values that are higher than the true value. concepts we need to understand in experimental error, **accuracy** and **precision**. Accuracy is how close your value or measurement is to the correct (true) value, and precision is how close repeated measurements are to each other. If the targets below represent attempts to hit the bulls eyes in an archery contest, they represent two types of error. The ones in the left image represents systematic error as all hits are to the left of the bulls eye. This kind of error would occur if you use an old ruler to measure length, but the ruler was worn down over time so it was no longer twelve inches. If you average all measurements that contain systematic error, you still miss the true value. On the right the holes are scattered around the bulls eye in relatively equal directions, so the error is random. If you average the random error, you actually get a good estimate of where the bulls eye is. So to compensate for systematic error you need to recognize it, and adjust for it. You can compensate for random error by making multiple measurements and averaging them.

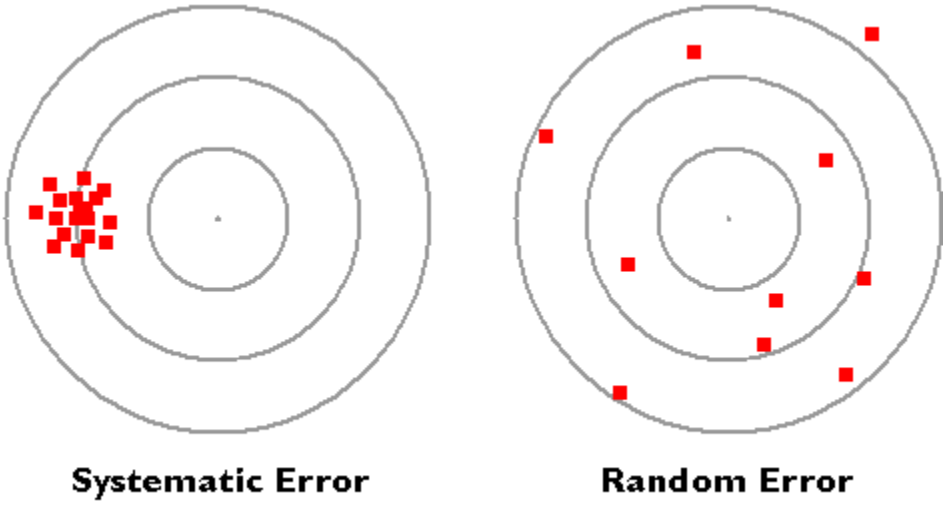


Figure 2.5.6: Systematic error has a bias where all measurements are offset, this is typically due to a flaw in an instrument, or how the operator uses it. Random error is evenly distributed around the true value.

Accuracy

The accuracy of a measurement is how close it is to the real value. If error is random, we can improve the accuracy by taking several measurements and using the average value. Thus, the average value of all the measurements on the right target of Figure 2.5.6 is very close to the center, but the average on the left side is not. We often use percent error to describe the accuracy of a measurement.

Percent Error

$$Percent\ Error = \frac{|Experimental\ Value - Theoretical\ Value|}{Theoretical\ Value}(100) \tag{2.5.5}$$

Where the theoretical value is the true value and the experimental value is the measured value. Note, some texts omit the absolute value sign, which means some measurements would have a positive percent error and others would have a negative. The problem

with that is if one wanted to know the average percent error for a series of random measurements, the positive and negative values would cancel and indicate a lower average value than is real. Note, that is a different question than what is the percent error of the average value, in which case you would calculate the average value, and base the error on that. The advantage of omitting the absolute value sign is a positive value means your reading is too high, and negative means it is too low. If you make multiple measurements, it is best to use the absolute value sign, and if only one, it does not really matter. It would probably be best to ask your instructor if you are in a lab class.

Precision

Precision is a measure of how close successive measurements are to each other. Precision is influenced by the scale, and when reporting a measurement, you report all certain values, and the the first uncertain one (which you "guesstimate"). Precision is how close successive measurements are and is often described by the Standard Deviation of a lot of measurements. This class will not do standard deviation counts, but you should quickly skim the chem 1402 section 1B2.121 Standard Deviation.

Note

For Random error you can average many values and get a pretty accurate reading of the true value, but for systematic error the average is off because all measurements had the same systematic error. Random error often results in the symmetric bell curve distribution where the frequency of over estimated values the same as the frequency of underestimated values. **Thus scientists often make a large number of measurements and use the average value, which only works for random error.**

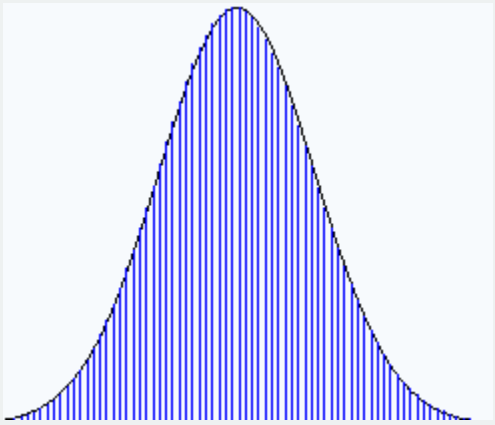


Figure 2.5.8: Random error distribution resulting in the symmetric bell curve.

Test Yourself

Homework: Section 2.5

Graded Assignment: Section 2.5

Query 2.5.1

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Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Stephen Lower (Simon Fraser University)
- Paul Flowers, Klaus Theopold & Richard Langley et al, General Chemistry at OpenStax CNX
- Vincent Belford (H5P interactive modules)

Answers

Exercise 2.5.1

Height = 0.4 or maybe 0.5 or 0.3, width = 0.2, although maybe 0.1 or 0.3, depending on how you look at it.

Exercise 2.5.2

Height = 35 or maybe 36 or 37, width = 22, although maybe 21 or 23, depending on how you look at it.

Exercise 2.5.3

Height = 35.2 or maybe 35.1 or 35.3, width = 21.3, although maybe 21.4 or 21.5, depending on how you look at it.

Exercise 2.5.4

Height = 35.31 or maybe 35.32 or 35.30, Width = 21.40, although maybe 21.41, depending on how you look at it.

Exercise 2.5.5

Two, you do not even need to do the math to answer this question.

Exercise 2.5.6

3, the answer is 100. or 1.00 x 10^2 because both numbers are known to the "ones" place.

Exercise 2.5.7

$$3 \times 21.3cm = 63.9cm \tag{2.5.6}$$

, the number has three significant digits because the three sheets of paper is an exact number and the 21.3cm has three significant digits.

Exercise 2.5.8

A

$$3 \times 35.2cm = 105.6cm = 106cm \tag{2.5.7}$$

, the number has three significant digits because the three sheets of paper is an exact number and the 35.2cm has three significant digits.

Exercise 2.5.9

$$63.9cm \times 105.6cm = 6747.84cm^2 = 6750cm^2 \tag{2.5.8}$$

Although we used four sig figs in the calculation for the height, both numbers really had 3 sig figs and so we round down to 3 sig. figs.(see Note below if you came up with 6770cm^2)

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2.6: Scientific Notation

Converting Floating Decimal Point Numbers to Scientific Notation

A closer look at scientific notation shows that the number is being represented by two parts, the pre-exponential and the exponential (10 to a power). To convert a floating decimal point number to scientific notation you need to multiply or divide the pre-exponential by 10 to a power, until there is just one digit to the left of the decimal point. Because this changes the value of the number, you need to divide or multiply that pre-exponential by 10 to a power (the second term of the notation, noting that division is represented by 10 to a negative exponent), so that the combined expression equals the value of the floating decimal point number.

This is easy to see by looking at the solutions to the next two examples, where you convert a fraction and a number greater than (or equal to) 10 to scientific notation.

Convert the following floating numbers to scientific notation

Exercise 2.6.3

0.00456 (A number smaller than 1)

Answer available at the end of this section.

Note, once you get the hang of this you can just count the number of digits you need to move the decimal to bring the original number to one digit to the left of the decimal and multiply that number by 10 to the negative value of those digits. Here we moved the decimal three positions to the right, so we multiply by 10^{-3} (ie., divide by 10^{3}).

Exercise 2.6.4

456.00, (A number larger than or equal to 10)

Answer available at the end of this section.

Note, once you get the hang of this you can just count the number of digits you need to move the decimal to bring the original number to one digit to the left of the decimal and multiply that number by 10 to the value of those digits. Here we moved the decimal two positions to the left, making it 100 times smaller, so we multiply by 10^{2} to maintain the same value.

Significant Digits:

How can you express the number 400 to 2 significant digits?

You must use scientific notation.

4.0×10^{2}

As scientific notation always has a decimal point there is never a problem expressing significant digits.

Multiplication and Division

Try plugging the following number into your calculator. Odds are you will get an error message, and you need to know how to solve calculations that are bigger or smaller than your calculator can handle.

$$\frac{\left(9.47 \times 10^{598}\right)\left(7.39 \times 10^{-98}\right)}{\left(1.432 \times 10^{-645}\right)\left(2.46 \times 10^{4}\right)} \tag{2.6.1}$$

The trick is to break the above equation into two parts, one dealing with the pre-exponential, and one dealing with the powers of 10. The following video shows how to do this, and the above problem is solved in the exercise 2.6.5 below the video.

Example 2.6.1: multiplication and division with scientific notation

Solve the following problem

$$\frac{\left(3.00 \times 10^{480}\right)\left(4.0 \times 10^{-20}\right)}{\left(2.00 \times 10^{-500}\right)\left(1.50 \times 10^{70}\right)}$$

Solution

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

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Video 2.6.1: Scientific notation in multiplication and division. https://youtu.be/nIFIZkkNRAk

Ans=4.0x10^{890}

Deeper Look into the Basics

Why do you add exponents when you multiply and subtract when you divide? A good way to understand this is to look at some simple operations you can do off your head. Like what is 100 times a 1000, or 1000 divided by 100, noting 100 = 10^2 and 1000 = 10^3.

a. One hundred times one thousand

$$100 \times 1000 = 100,000 \tag{2.6.2}$$

$$10^2 \times 10^3 = 10^5$$

$$10^{(2+3)} = 10^5$$

b. one thousand divided by 100

$$\frac{1000}{100} = 10 \tag{2.6.3}$$

$$\frac{10^3}{10^2} = 10$$

$$10^3 \times 10^{-2} = 10$$

$$10^{(3-2)} = 10$$

$$10^1 = 10$$

Exercise 2.6.5

Solve the following problem.

$$\frac{\left(9.47 \times 10^{598}\right)\left(7.39 \times 10^{-98}\right)}{\left(1.432 \times 10^{-645}\right)\left(2.46 \times 10^{4}\right)}$$

Answer available at the end of this section.

Addition and Subtraction

This can easily be solved with your calculator, the challenge is how do you determine the number of significant Figures in the answer. To do this you need to line them up by multiplying everything by the same power of 10.

Consider the following problem

$$4.860 \times 10^{12} + 9.7 \times 10^{10} + 3.68 \times 10^{11}$$

Would you believe the answer has four significant Figures? (The solution to this problem is below the following video.)

Example 2.6.2: Addition and subtraction with scientific notation

Solve the following problem and report answer to the correct number of significant Figures:

$3 \times 10^{23} + 3.00 \times 10^{25}$

Solution

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

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Video 2.6.2: Scientific Notation in addition and subtraction. https://youtu.be/Ghe32OnScdI

(3.00×10^{25})

Exercise 2.6.4

$4.860 \times 10^{12} + 9.7 \times 10^{10} + 3.68 \times 10^{11}$

Answer available at the end of this section.

Test Yourself

Homework: Section 2.6

Graded Assignment: Section 2.6

Activity 2.6.1

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NOTE: For worked out Solutions of Activity 2.6.2 go to homework section 2.7

The next activity deals with addition and subtraction of numbers in scientific notation.

Activity 2.6.2

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Answers

Exercise 2.6.3

Step 1: Since this number is less than 1 you need identify the power of 10 which you can multiply it by and set one digit to the left of the decimal. Here, we can multiply by 1000 or 10^3 and force the original number to have one digit to the left of the decimal, but we have changed its value.

0.00456 x 10^3 = 4.56 (note, which is not equal to the original number)

Step 2: Since we multiplied it by 10^3 we must also divide by 10^3 (effectively multiplying the original number by 1 and not changing it's value). Than we express the factor in the denominator as a power of 10.

$$(0.00456) = (0.00456)\left(\frac{10^3}{10^3}\right) = \left(\frac{0.00456 \times 10^3}{10^3}\right) = \left(\frac{4.56}{10^3}\right) = 4.56 \times 10^{-3} \tag{2.6.4}$$

Exercise 2.6.4

Since this number has more than one digit to the left of the decimal, we must divide it by a factor of 10 to make it smaller so that one digit is to the left of the decimal. Then we must multiply everything by 10 to the same power that we divided by 10 in the first step.

$$(456.00) = (456.00)\left(\frac{10^2}{10^2}\right) = \left(\frac{456.00}{10^2}\right) \times 10^2 = 4.56 \times 10^2 \tag{2.6.5}$$

Exercise 2.6.5

$$= \left(\frac{9.47x7.39}{1.432x2.46}\right)\left(\frac{10^{598} \times 10^{-98}}{10^{-645} \times 10^4}\right)$$
$$= \left(19.86 \times 10^{(598-98+645-4)}\right)$$
$$= \left(19.9 \times 10^{1141}\right) = \left(1.99 \times 10^{1142}\right)$$

Closer

Exercise 2.6.4

You first need to express all numbers to the same power so you can line up the decimal point. It is suggested that you choose the largest power and make everything else a fraction.

4.860 x 10^{12}

+0.097 x 10^{12}

+0.368 x 10^{12}

5.325 x 10^{12}

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2.7: SI Prefixes

Introduction

We have all heard of SI prefixes, for example you may have a 16 Gbyte flash drive, where the G stands for giga and means 10^9, and that is a much more convienent way of expressing the number of bytes than using the floating number of 16,000,000,000 bytes. Chemists need to be able to convert between different ways of expressing, but the first step is to memorize the unit definitions themselves. It is important that you do this as you will not be able to follow the science later on in the course if you do not understand the meanings of the terms used. It is like trying to write a sentence without knowing the alphabet What you notice is that most of the units you are required to memorize are multiples of 10^3, and if you look at the above 16 Gbyte flash drive, when we wrote it as a floating decimal point number we placed a comma between every "000", which is 10^3,moved it up or down 1000 positions, so Giga (10^9), had 9 zeros with 3 commas. The result is you can express any number with between 1 and 999 values to the left of the decimal, followed by the appropriate SI prefix. So we would say 16 Gigabytes and not 16,000 Mbytes of 16,000,000 kbytes.

The convention of using an SI prefix with a number between 1 and 999 is often called engineering notation and is a common way of representing very large and small numbers.

SI Prefixes

SI prefixs are commonly used to represent very large and small numbers. You need to memorize many of these and know how to convert between them.

*Table 2.7.1: SI prefixes. Note **Chem 1300 students need to memorize all the values in red**.*

Multiplier	Name	Abbreviation		Name	Multiplier
10^{+30}	quetta*	Q	q	quecto*	10^{-30}
10^{+27}	ronna*	R	r	ronto*	10^{-27}
10^{+24}	yotta	Y	y	yocto	10^{-24}
10^{+21}	zetta	Z	z	zepto	10^{-21}
10^{+18}	exa	E	a	atto	10^{-18}
10^{+15}	**peta**	**P**	**f**	**femto**	**10^{-15}**
10^{+12}	**tera**	**T**	**p**	**pico**	**10^{-12}**
10^{+9}	**giga**	**G**	**n**	**nano**	**10^{-9}**
10^{+6}	**mega**	**M**	**m**	**micro**	**10^{-6}**
10^{+3}	**kilo**	**k**	**m**	**milli**	**10^{-3}**
10^{+2}	**hecto**	**h**	**c**	**centi**	**10^{-2}**
10^{+1}	**deca**	**da**	**d**	**deci**	**10^{-1}**

*quetta, ronna, ronto & quecto were added in 2022 (https://www.bipm.org/en/cgpm-2022/resolution-3).

NOTE: The Angstrom is a common unit of length, 1 Å = 10^{-10}m). It is not an SI prefix, but is a common unit of length that is used when describing bond lengths and distances between atoms (and ions) in a crystal.

Use of SI prefixes

For real large or small numbers, the convention is to place a number with between 1 and 999 in front of the SI prefix, and use the appropriate prefix to show the value. So a memory stick with 1,200,000 bytes would be written as 1.2 Mbyte, not as 1,200 kbytes.

In a later section you are going to have to learn how to convert between SI units, or even between SI units and scientific notation.

SI prefixes and Calculators

What is the ENG key on scientific calculators?

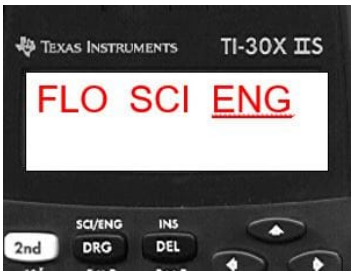


Figure 2.7.8: Three display modes on TI 30 calculator.

In Scientific Notation there is one digit to the left of the decimal, in Engineering notation there are 1, 2 or 3 digits to the left of the decimal, and you then use one of the Si prefixes that is of the magnitude 10^{3n}.

$$\begin{aligned} 1,330,000,000m &= 1.33 \times 10^{9}m = 1.33\ gigameters \\ 13,300,000,000m &= 13.3 \times 10^{9}m = 13.3\ gigameters \\ 133,000,000,000m &= 133 \times 10^{9}m = \ 133\ gigameters \end{aligned} \tag{2.7.1}$$

The ENG key takes large or small numbers and expresses them as integer multiples of 10^{3} or 10^{-3}, effectively converting them to values that can be expressed with SI prefixes. If you type 1,200,000 into your calculator and press ENG, it gives 1.2 x 10^{6}, so 1,200,000 bytes of memory on a memory stick would be 1.2 MBytes (you do not say 1,200 kBytes). If you try 0.12 and press ENG, it gives 120 x 10^{-3}, so 0.12g is 120 mg. So the ENG function quickly converts numbers to values that can easily be expressed with SI prefixes.

Test Yourself

Homework: Section 2.7

The following activities are drill and practice. You need to memorize the red SI prefixes in table 2.7.1. Note, on the flash cards you can skip a card by clicking

"I got it right" and can skip the ones you know.

Activity 2.7.1

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

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Flash cards that go from SI prefix name to value. *NOTE: If you already know an answer you can just click "I got it right" and move*

on to the next. Everytime you reload these should rearrange.

Activity 2.7.2: Prefix to Value

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Flash cards that go from SI prefix value to name. *NOTE: If you already know an answer you can just click "I got it right" and move on to the next. Everytime you reload these should rearrange.*

Activity 2.7.3: Value to Prefix

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Flash cards that go from SI prefix symbol to value. *NOTE: If you already know an answer you can just click "I got it right" and move on to the next. Everytime you reload these should rearrange.*

Activity 2.7.4: Abbreviation to Prefix

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Contributors and Acknowledgements

- Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Elena Lisitsyna & Vincent Belford (H5P interactive modules)

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2.8: Unit Conversions and Dimensional Analysis

Equivalence Statements

As chemists have different ways of expressing measurements they need to be able to convert between different units. Central to this is the concept of an equivalence statement which says two ways of representing the same thing are equivalent. For example 12 in = 1 foot is an equivalence statement. An equivalence statement allows you to convert from one unit to the other, and this is done by creating a conversion factor. A conversion factor is simply the ration of one part of the equivalence statement to the other, where the numerator has the unit you want to convert to, and the denominator has the unit you want to transform from.

12 in = 1 foot

has two conversion factors

1. to convert from feet to inches multiply by the conversion factor of

$$\left(\frac{12\ in}{1\ foot}\right) \tag{2.8.1}$$

2. to convert from inches to feet multiply by the conversion factor of

$$\left(\frac{1\ foot}{12\ inches}\right) \tag{2.8.2}$$

There are three fundamental types of conversions that you will need to be able to perform. The first involve different ways of expressing numbers within the same scale, which include transforming floating numbers to SI prefixes and scientific notation. The second involves converting from one scale to another, like the metric system to the English [read: American] system, and you are expected to know that 2.54 cm=1 inch [read: memorize that]. The third involves physical constants of specific substance, where you convert from one type of physical measurement to another. Density is such a physical constant, and it allows you to convert mass measurements to volume.

What is consistent with each of these three different types is that they all represent an equivalence statement, and from the equivalence statement you can generate a conversion factor.

Three Basic Types Conversions

A. Different prefixes of same base scale (cm -> Km)
B. Conversions between scales (cm -> in)
C. Physical Constants d=M/V

Conversions Involving SI Prefixes

2 Basic Techniques:

1. Basic Technique : Use two equivalence statements
2. Advanced Technique: Use one equivalent statement

Basic Technique

How many micrograms are in a kilogram?

Set up conversion factors based on an equivalence statements.

1 kg = 10^3 g and 1 µg = 10^{-6} g

so, this gives two conversion factors

$$\left(\frac{10^3 g}{1 kg}\right) and \left(\frac{1 \mu g}{10^{-6} g}\right) \tag{2.8.3}$$

You have two conversion factors. In the first you will convert kilograms to grams, and in the second you will convert the grams to micrograms.

$$(1 kg)\left(\frac{10^3 g}{1 kg}\right)\left(\frac{1 \mu g}{10^{-6} g}\right) = 10^{3-(-6)} \mu g = 10^9 \mu g \tag{2.8.4}$$

In the advanced technique (below) you will solve this problem with only one conversion factor.

Example 2.8.1

$3x10^{23}$fsec = ? Gsec

Solution

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View this YouTube Video online

Video 2.8.1: Basic technique for unit conversions. https://youtu.be/Jlfk7XjDnwM

($3x10^{-1}$ Gsec)

Exercise 2.8.1

Convert 32.4 ng to Mg.

Answer available at the end of this section.

Advanced Technique

How many micrograms are in a kilogram?

Set up a single conversion factor based on an equivalence statement.

The trick here is to multiple both SI prefixes by the number that makes them equal to one. That is, if a micro is one millionth, than there are a million micro in one, and and likewise if a kilo is a thousand, than a thousandth of a kilo is one, and so a million micrograms equals a thousandth of a kilogram.

$$1 \mu g = 10^{-6} g \therefore 1 g = 10^6 \mu g \tag{2.8.5}$$

$$1 kg = 10^3 g \therefore 1 g = 10^{-3} kg \tag{2.8.6}$$

so, this gives two conversion factors

$$\frac{10^6 \mu g}{10^{-3} kg} = \underbrace{\frac{10^9 \mu g}{1 kg}}_{\text{converts kg to } \mu g} \leftarrow or \rightarrow \frac{10^{-3} kg}{10^6 \mu g} = \underbrace{\frac{10^{-9} kg}{1 \mu g}}_{converts\ \mu g\ to\ kg} \tag{2.8.7}$$

(Use the ratio which has the starting unit in the denominator and ending unit in the numerator.)

See Video 2.8.2 for the advanced technique combined with scientific notation
(next video below)

Exercise 2.8.2

Convert 32.4 ng to Mg..

Answer available at the end of this section.

SI Prefix Conversions and Scientific Notation

The following video shows you how to do SI conversions of numbers expressed in scientific notation. The trick is to write the SI prefix in terms of its power of 10.

Example 2.8.2

Convert: 3 x 10^{23} kiloseconds to nanoseconds.

Solution

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 2.8.2: Advanced technique with SI notation. https://youtu.be/43wWTLB7GS0

(3 x 10^{35} nsec)

Exercise 2.8.3

Convert 3.57x10^{14} ml to ml

Answer available at the end of this section.

Conversions Between Different Scales

There are typically two different types of conversions between scales; those between different systems (like inches to meters), and those between SI base units and derived units (like ml to m^3). America still uses the English or "British Imperial System" for common measurements like mass, length and volume, even though the British don't, and thus we often need to convert between the two different systems. What is required, as in any conversion, is an **equivalence statement**. In this class you are only responsible to memorize two conversion factors of this nature

You need equivalence statements between different units. These are the same techniques as the previous, just that the conversions are between different types of units.

Note

UALR students are required to memorize the following conversion factors

- 2.54 cm=1 in (exact) (Metric length to English Length)
- 1 ml = 1 cm^3 (Metric volume to Metric length cubed)

Exercise 2.8.4

How many dm are in an object which is 2.450 feet long?

given 12 in = 1 ft,

Answer available at the end of this section.

Example 2.8.3

What is the volume in Gallons of a 1.00 cubic foot object?

(Try this first, the answer is at the bottom of the video.)

Solution

Note, you are required to know a 1 ml =1 cm^3, 1 in = 2.54 cm, and of course the SI prefix conversions.

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 2.8.3: Conversions between scales

(7.48 gal)

Exercise 2.8.5

How many gallons are in a cuboid can that is 2.05 ft wide, 12.4 inches long and 1.30 yards deep?

Given: 4 qt = 1 gal, 1.057qt=1L, 12in = 1 ft, 3 ft = 1 yd:

Answer available at the end of this section.

Conversions Involving Physical Constants

Many substances have physical properties showing a linear relationship that can be described by the equation of a straight line Y=mX+b. If these are extensive properties (section 1.6.1), then the Y intercept (b) equals zero and this takes the form of Y=mX, which is an equivalence statement, and allows one to convert measurements in the unit of X to those of the unit of Y. These are often tabulated by the constant m $\left(m = \frac{Y}{X}\right)$, which allows conversions of measurements between the units of X and Y.

This can be understood by looking at one such constant, density, which is the linear proportionality constant for the extensive

properties of mass and volume, where m = dv (the y intercept is zero because if you have nothing, its mass and volume both equal zero - review extensive properties). So if you know the density of something, you can convert its mass to volume, or its volume to mass. Therefor it is common to compile tables of densities and other physical constants.

It should be stated that although these generate conversion factors, they are really describing relationships, that is, mass and volume describe completely different aspects of matter and density allows you to convert from one to another. This is completely different than definind 12 inches as a foot, where inches and feet are different units that describe the same thing.

Conversions Involving Density

Density is defined as

$$d = \frac{m}{v} \tag{2.8.8}$$

This gives the equivalence statement

$$m = dv \tag{2.8.9}$$

and there are thus two conversion factors

$$\underbrace{\frac{m}{d}}_{\text{converts mass to volume}} \leftarrow or \rightarrow \underbrace{dv}_{\text{converts volume to mass}} \tag{2.8.10}$$

That is, dividing the mass by the density converts it to its volume, and multiply the volume by the density converts its volume to its mass

Example 2.8.4

How many liters is the volume of 12.140 kg of Iridium? Use 22.65 g/ml as the density of Iridium

(Please do this before looking up the answer at the bottom of the video.)

Solution

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 2.8.4: Conversion Factors involving physical scales.

0.5360 L, note, the number of sig figs was determined by the density and not the starting mass.

Exercise 2.8.6

How many pounds would 61.8 gallons of water weight if at a given temperature the water has a density of 0.98g/mL?

Answer available at the end of this section.

Note

If you look at the answer of exercise 2.8.6 you will see that each **equivalence statement is in a set of parenthesis**. You will also note that all three types of conversion factors are used. ($\frac{4qt}{1gal}$) is between two different units (4 quarts = 1 gallon), ($\frac{1000mL}{L}$) is between different SI prefixes (1000mL=1L) and ($\frac{0.98g}{ml}$) is a physical constant, the density of water (0.98g H_2O = 1mL H_2O, with two sig figs).

Test Yourself

Homework: Section 2.8

Graded Assignment: Section 2.8

Query 2.8.1

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Answers

Exercise 2.8.1

$$(32.4ng)\left(\frac{10^{-9}g}{ng}\right)\left(\frac{Mg}{10^{6}g}\right) = 32.4\times10^{(-9-6)}Mg = 32.4\times10^{-15}Mg = 3.24\times10^{-14}Mg \tag{2.8.11}$$

Exercise 2.8.2

1 g = 10^{-6} Mg = 10^{9} ng

$$(32.4ng)\left(\frac{10^{-6}Mg}{10^{9}ng}\right) = \left(32.4\times10^{(-6-9)}Mg\right) = \left(32.4\times10^{-15}Mg\right) = \left(3.24\times10^{-14}Mg\right) \tag{2.8.12}$$

Exercise 2.8.3

$$\left(3.57\times10^{14}\mu l\right)\left(\frac{10^{3}ml}{10^{6}\mu l}\right) = \left(3.57\times10^{(14+3-6)}ml\right) = \left(3.57\times10^{11}ml\right) \tag{2.8.13}$$

Exercise 2.8.4

$$(2.450\,ft)\left(\frac{12\in}{1\,ft}\right)\left(\frac{2.54\,cm}{1\,in}\right)\left(\frac{10^{1}dm}{10^{2}cm}\right) = (7.4676\,dm) = (7.468\,dm) \tag{2.8.14}$$

Exercise 2.8.5

$$(2.05ft)\,(12.4in)\,(1.3yd)\left(\frac{3ft}{1yd}\right)\left(\frac{12in}{1ft}\right)^{2}\left(\frac{2.54cm}{1in}\right)^{3}\left(\frac{1mL}{1cm^{3}}\right)\left(\frac{1L}{1000mL}\right)\left(\frac{1.057qt}{1L}\right)\left(\frac{1gal}{4qt}\right) = 61.8gal \tag{2.8.15}$$

Exercise 2.8.6

$$(61.8gal)\left(\frac{4qt}{1gal}\right)\left(\frac{1L}{1.057qt}\right)\left(\frac{1000ml}{1L}\right)\left(\frac{0.98g}{mL}\right)\left(\frac{1lb}{453.6g}\right) = 505.27lb = 510lb \tag{2.8.16}$$

Note the answer has 2 significant digits because of the density.

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2.9: Fractions and Percent

Introduction

Fractions and percent essentially tell you the same thing, they describe the ratio of a part of the whole.

Fraction

$$Fraction = \frac{Part}{Whole} = \frac{Part}{\sum Parts} \tag{2.9.1}$$
$$\left(\sum \text{ means sum of}\right)$$

Exercise 2.9.1

Consider Mixing 23.5 g of table salt (NaCl) with 91 g of water

a. What is the fraction salt?
b. What is the fraction water?

Answer a

$$\text{fraction salt} = \frac{m_{salt}}{m_{salt} + m_{water}} = \frac{23.5g}{23.5g + 91g} = \frac{23.5g}{114.5g} = 0.20524 = 0.205$$

Answer b

$$\text{fraction water} = \frac{m_{water}}{m_{salt} + m_{water}} = \frac{91g}{23.5g + 91g} = \frac{91g}{114.5g} = 0.79476 = 0.79 \tag{2.9.2}$$

Note, the number of significant figures in answer (a) above is different than answer (b).

The sum of the fractions equals 1, that is all the parts combine into the whole, and 1 is a defined number.

$$fraction_{salt} + fraction_{water} = \frac{m_{salt}}{m_{salt} + m_{water}} + \frac{m_{water}}{m_{salt} + m_{water}} \tag{2.9.3}$$

$$= \frac{m_{salt} + m_{water}}{m_{salt} + m_{water}}$$

$$= 1 \tag{2.9.4}$$

In general this can be expressed as the summation of each fraction starting with i=1, being summed up to i=n, when there are n parts (fractions)

$$\sum_{i=1}^{n} (fraction)_i = 1 \tag{2.9.5}$$

So there was a second way to solve exercise 1b, if we know the answer to 1a

$$\begin{aligned} fraction_{salt} + fraction_{water} &= 1 \\ \therefore \\ fraction_{water} &= 1 - fraction_{salt} \\ &= 1 - 0.205 \\ &= 0.795 \end{aligned} \tag{2.9.6}$$

Note

The number of significant digits in solving problems can vary depending on the mathematical technique you use

Percentages

Percentages are 100 times the fraction, and so we can redo all of the above material in terms of percents. Note, the 100 is a defined number and does not contribute to the significant figures (100% of a sample is defined as the entire sample). The **symbol for percent is %**.

$$Percent = Fraction(100) = \frac{Part}{Whole}(100) \tag{2.9.7}$$

Exercise 2.9.2

Consider Mixing 23.5 g of table salt (NaCl) with 91 g of water

a. What is the percent salt?
b. What is the percent water?

Answer a

$$\%\ salt = \frac{m_{salt}}{m_{salt} + m_{water}}(100) = \frac{23.5g}{23.5g + 91g}(100) = \frac{23.5g}{114.5g}(100) = 20.524\% = 20.5\% \tag{2.9.8}$$

Answer b

$$\%\ water = \frac{m_{water}}{m_{salt} + m_{water}}(100) = \frac{91g}{23.5g + 91g}(100) = \frac{91g}{114.5g}(100) = 79.476\% = 79\% \tag{2.9.9}$$

Note, the number of significant figures in answer (a) above is different than answer (b).

The sum of the percentages equals 100, that is all the parts combine into 100%.

$$\%_{salt} + \%_{water} = \frac{m_{salt}}{m_{salt} + m_{water}}(100) + \frac{m_{water}}{m_{salt} + m_{water}}(100) \tag{2.9.10}$$

$$= \frac{m_{salt} + m_{water}}{m_{salt} + m_{water}}(100)$$

$$= 1(100) = 100\% \tag{2.9.11}$$

In general this can be expressed as the summation of each fraction starting with i=1, being summed up to i=n, when there are n parts (fractions)

$$\sum_{i=1}^{n} (percentage)_i = 100\% \tag{2.9.12}$$

So there was a second way to solve exercise 1b, if we know the answer to 1a

$$\begin{aligned} \%_{salt} + \%_{water} &= 100\% \\ &\therefore \\ \%_water &= 100\% - \%_{salt} \\ &= 100\% - 20.5\% \\ &= 79.5\% \end{aligned} \tag{2.9.13}$$

Exercise 2.9.3

What mass of Water do you need to add to 55.4g of salt to make a solution that is 15.0% Salt?

Hint: Solve this in terms of fractions, where fraction of salt = $\frac{\%salt}{100}$

Answer available at the end of this section.

Fractions & Percent are Unitless

Lets look at a bag with of 100 nickels and 100 dimes, and measure the mass on two different scales and then calculate the percent of the mass in the bag that is dues to the nickles and the mass that is dimes.

	Mass in grams	mass in ounces
100 nickels	500.0 g	17.63 oz
100 dimes	226.8 g	8.000 oz
total mass	726.8 g	25.63

In units of grams

$$\%nickle = \frac{500g}{726.8g}(100) = 68.79\% \quad \%dimes = 100\% - 68.79 \tag{2.9.14}$$

In units of ounces

$$\%nickle = \frac{17.63oz}{25.63oz}(100) = 68.79\% \quad \%dimes = 100\% - 68.79 \tag{2.9.15}$$

Note, both grams and ounces are units are measurements of mass, and the mass percent is independent of the unit. But if you are measuring the number of entities, the above bag is 50% nickels and 50% dimes.

Test Yourself

Homework: Section 2.9

Interactive Quiz 2.9.1

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View this Interactive Content online

Solutions to the above problems can be found in the homework section 2.9.

Answers

Exercise 2.9.3

f_s=fraction salt, m_s=mass salt, m_w=mass water, if percent salt = 15.0%, the f_s= 0.150

$$f_s = \frac{m_s}{m_s + m_w} \tag{2.9.16}$$

$$\text{solve for } m_w$$

$$f_s\left(m_s + m_w\right) = m_s$$

$$f_s m_s + f_s m_w = m_s$$

$$f_s m_w = m_s - f_s m_s = m_s\left(1 - f_s\right)$$

$$m_w = \frac{(1 - f_s)}{f_s} m_s = \frac{(1 - .150)}{.150} 55.4g = \frac{850}{.150} 55.4g = 314g$$

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2.10: GRAPHING AND TEMPERATURE CONVERSIONS

Ancillary Material

Graph Paper (PDF of blank sheet with a graph grid that can be printed)

Graph Paper (html that can be used with iPad)

Functional Dependencies

A graph can be used to show the relationship between two related values, the dependent and the independent variables.

- **Dependent Variable:** A measurable variable whose value depends on the independent variable. Often represented by the algebraic symbol y.
- **Independent Variable:** A measurable variable of which when changed, the value of the dependent variable changes. Often represented by the algebraic symbol x.

$$y = \text{function of x} \quad or \quad y = f(x) \tag{2.10.1}$$

Such a relationship would be the maximum amount of table salt (sodium chloride) you can dissolve in water as a function of temperature, which is represented by the data in table 1B.5.1. This represents a saturated solution, and what you notice is that if you raise the temperature of the water more table salt can dissolve. That is, if you have 26.1 g salt at 20.0 ^{o}C dissolved in 100 grams of water (which is the maximum amount that can dissolve in 100 grams of water), and add another 2 grams, it just falls to the bottom of the solution and the temperature does not change. If you then raise the temperature the solid at the bottom starts to dissolve and somewhere around 70 degrees all the salt dissolves. So the solubility is the dependent variable, and its value depends on the temperature, the independent variable. You noted in the above example that changing the dependent variable (solubility) had no effect on the temperature (independent variable).

In this case Equation 1B.5.1 could be expressed as

$$Solubility = f(Temperature) \tag{2.10.2}$$

Table 2.10.1: Solubility of table salt as a function of water. In this graph 6 measurements of the solubility were made at different temperatures. Here, the solubility is the maximum amount of salt that can be dissolved in 100 grams of water, and any extra salt beyond that falls to the bottom of the container as a precipitate (solid).

Data Measurement number	Temperature (^{o}C)	Solubility (g NaCl/100 g water)
1	0.00	35.2
2	20.0	36.1
3	40.0	37.0
4	60.0	37.7
5	80.0	38.4
6	100.0	39.0

A plot can then be made of the data where the dependent variable (y) is plotted on the vertical axis (ordinate axis) and the independent variable(x) is plotted on the horizontal axis. Figure 1B.5.1 represents a plot of the data from table 1B.5.1, where each point represents one of the data measurements, with the values typically expressed in terms of (x,y). So the third data point would have values of (60.0,37.0) representing the temperature and solubility.

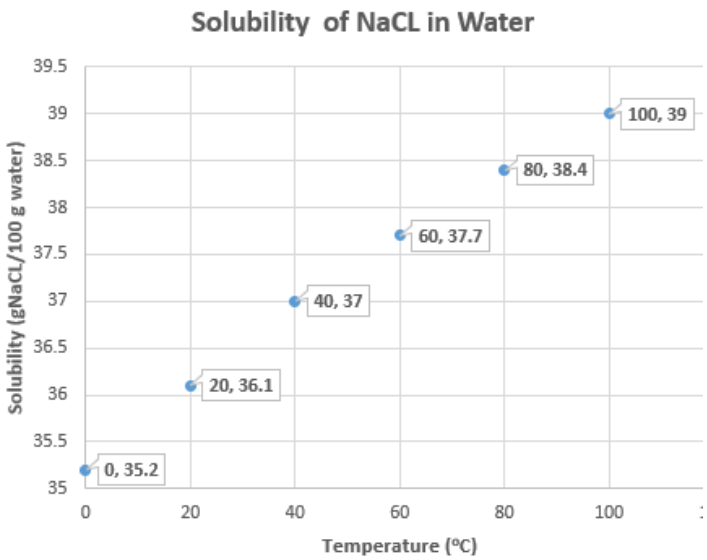


Figure 2.10.1: Data from table 1B.5.1 plotted with temperature along the x-axis and solubility along the y Axis

It should be noted that not all plots involve dependent and independent variables, and that sometimes both variables are dependent on another entity, which is not being plotted. This is the case for temperature conversions from Fahrenheit to Celsius, where both scales are measuring an intensive property, the temperature of some object or substance, and it is the temperature that changing. Or a plot of mass vs. volume for a pure incompressible substance, which are both extensive properties that depend on the amount of substance being measured.

Graphing Linear Functions

Equation 1B.5.3 is a linear equation

$$y = mx + b$$

- m is called the slope of the line, which is the ratio of the change in y (the rise) to the change in x (the run). Since the line is linear, the slope is the same between any two points along the line.
- b is called the y-intercept, which is the value of y when x=0.

Steps for Graphing a Linear Function

1. Pick your range and scale: In graphing a linear function the first step is to identify the range and scale for each axes. The range is defined by the minimum and maximum value along each of the axes. The scale is defined by the values of each line in the grid, which is used both to plot the data points, and once graphed, to read values from the graph. There are often minor and major marks given to the grid, where the major marks show the numerical value of that position.

    - Choosing the Range: There are two approaches. The first is to use the highest and lowest data value as the lower and upper bound of the range, although you may want the actual bounds to be some easy number to read, like a multiple of ten. The second approach is to look at the data, and see if the data is bounded by some real limit. For example, in the above data we are looking at the solubility of a salt in liquid water, so the temperature range may want to be defined by the temperature range where water is a liquid, even if your highest data point is not at the boiling point, and your lowest is not at the freezing point.
    - Choosing the Scale: Here you want the grid line to be easy to read. If possible, multiple's of 10 are very useful, but you need to look at the data. Want you want to avoid is where each mark is some hard to read value.

2. Plot data: Mark each data point on the graph (like in Figure 1B.5.1, but do not write down the x,y coordinate values, just mark a point on the graph to indicate where there is a data value.
3. Hand draw a best fit line that has an even distribution of points on both sides of the line. If the data is truly a linear function, there will be a random distribution of points along the line. Figure 1b.5.3 shows such a "best fit line".

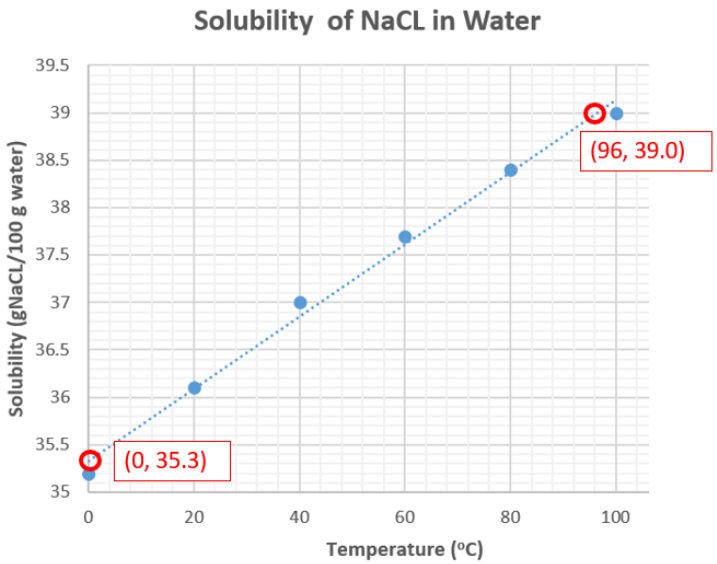


Figure \({2[.]10{.}\eatSpaces{3} \endSpaces}]): Best fit line that evenly describes the data. Not we do not connect the points, and some of the points are not on the line.

4. Pick two points to calculate the slope from. Start at the left side and find the point on the line with the smallest values where the line crosses the grid lines, here it is point (0,25.50), and then starting at the right end pick the point with the highest values (96,39.00).

    - Note the precision is defined by both the data and the scale on the grid lines.
        - For the y-axis the precision is defined by the data, and although we have a mark every 0.1 values, so it looks like we could report solubility to the .01 value, the original data was only known to the 0.1 value (table 1B.5.1). So so even though the scale on the graph is more precise, we can only use that precision to read from the scale, and not report the value. That is from the scale we could read to a value of 35.30 and not 35.3, but we must use 35.3 in our calculation
        - For the x-axis the original data has a precision to the 0.1 digit, that is, we have values like 20.0 ^{o}C. But the smallest value on the minor grid axis is 4 ^{o}C, and so we have to "guess" between them to report the uncertainty, and thus choose the precision to be to the "ones" position.
    - Note, if we choose points near to each other, we can lose significant digits. For example, if the second point was (4,35.5) the change in temperature would be 4-0 or 4 degrees (1 significant digit), while using the value on the graph give 96-0, or 96 degrees, which has two significant Figures.
    - That is, we report all certain, and the first uncertain data point, as read off the graph. A graph scale should never have a smaller scale than the raw data, that is, it can't add to the precision

5. Now calculate the slope

$$\text{slope} = m = \frac{\text{rise}}{\text{run}} = \frac{(y_2 - y_1)}{(x_2 - x_1)} = \frac{(S_2 - S_1)}{(T_2 - T_1)} = \frac{(39.0 - 35.3)}{(96 - 0)} = 0.039\ \frac{g\ NaCl}{100 g H_2O \cdot\ ^{o}C} \tag{2.10.3}$$

6. Now calculate the y intercept. Since the graph goes through zero, you can just read it from the graph

    - Since the x-axis includes zero, you can simply read it from the graph, b=35.3(gNaCl/100g water).

Note

If your graph does not include the zero point, you can extrapolate backwards from a given value.

We know that:

$$m = \frac{\Delta y}{\Delta x} = \frac{\Delta S}{\Delta T} \tag{2.10.4}$$

If we define S_{20}=solubility when T=20, and S_0=solubility when T=0, then b= S_0, and S_{20} comes from the graph on Figure 1B.5.3 where, **S_{20}=36.1(gNaCl/100g water)**.

$$m = \frac{S_{20} - S_0}{20 - 0} = \frac{S_{20} - b}{\Delta T_{(=20)}} \tag{2.10.5}$$

Where $\Delta T_{(=20)}$ means the value of ΔT= 20,

solving for b and including units in numerical calculation gives:

$$b = S_{20} - m(\Delta T_{(=20)}) = 36.1 \left(\frac{g\ NaCl}{100 g H_2O} \right) - 0.039 \frac{g\ NaCl}{100 g H_2O \cdot\ ^oC} (20\ ^oC) = 35.3 \left(\frac{g\ NaCl}{100 g H_2O} \right) \tag{2.10.6}$$

Temperature Conversions

Figure 1B.3.2 shows the relationship between the three commons scales with values given for absolute zero and the boiling and freezing points of water.

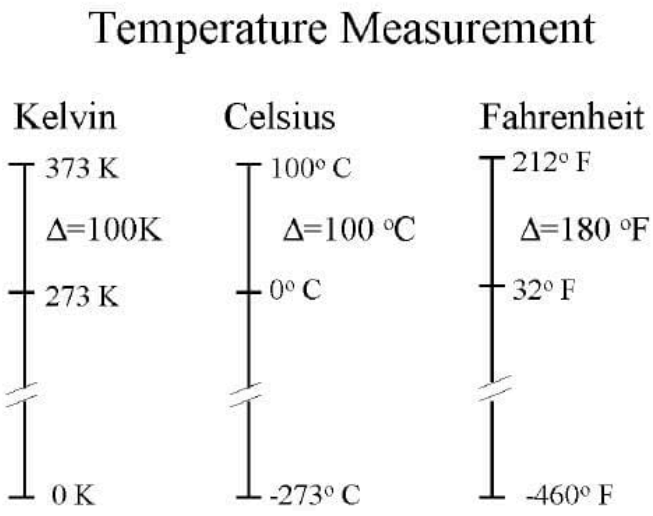


Figure 2.10.2: Relationship of the three common temperature scales

Celsius/Kelvin Conversions

Because the kelvin defines the Celsius scale, the conversion is simple.

Celsius to kelvin

$$T_{kelvin} = T_{celsius} + 273.15 \tag{2.10.7}$$

Kelvin to Celsius

$$T_{celsius} = T_{kelvin} - 273.15 \tag{2.10.8}$$

Celsius/Fahrenheit

Both Celsius and Fahrenheit are linear scales and a plot of data from the freezing point to the boiling point of water can show the

relationship. In Figure 1B.3.3 we plot the Fahrenheit values along the ordinate (y-axis) and Celsius along the abscissa (x-axis) and we note there are multiple ways of writing the equation relating Celsius and Fahrenheit.

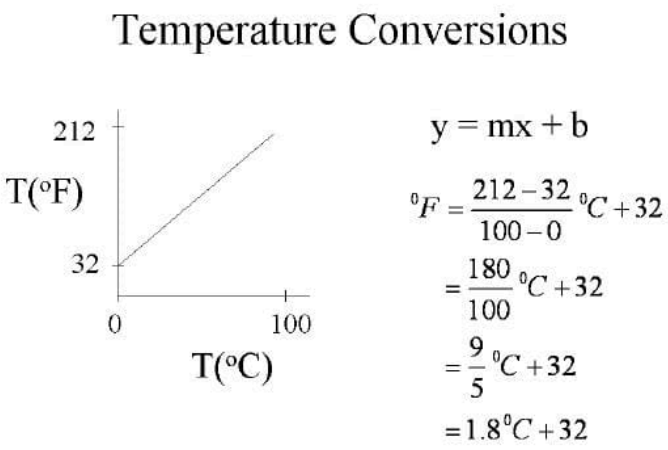


Figure 2.10.3: plot of T(F) as a function of T(C)

We note that the above equation has two forms

Celsius to Fahrenheit

$$T_F = 1.8T_C + 32 \tag{2.10.9}$$

Fahrenheit to Celsius

$$T_C = \frac{T_F - 32}{1.8} \tag{2.10.10}$$

These equations are much more complicated than the density equation because the y-intercept is not zero. In the following videos we will develop a quick technique for converting between these temperature scales called the Plus 40/minus 40 technique. So first, lets look at the difference between the linear plots of Fahrenheit vs. Celsius and mass vs. volume.

YouTube Video

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View this YouTube Video online

Video 2.10.1: Graphing Part 1 - In this video we compare the plot of Fahrenheit and Celsius to that of mass and volume, and see that the density plot is simpler to use because the y-intercept is zero

So what we do is find the temperature at which $T_F = T_C$ and make that value equal to zero

Exercise 2.10.1

At what temperature does $T_F = T_C$?

Answer available at the end of this section.

So the scales converge at -40 degrees.

Plus Forty/Minus Forty Technique

The strategy of the plus forty/minus forty techniques is to make the value where the two scales intersect equal to forty, then the y-intercept is zero, and so you can just multiple or divide by the slope. This is a three step process

1. Add 40 to the number (since the intersected at -40, that value is now zero)
2. multiple or divide by the slope
   - If going from Fahrenheit to Celsius, divide by 1.8
   - If going from Celsius to Fahrenheit, multiply by 1.8
3. Subtract the 40If going from Fahrenheit to Celsius, divide by 1.8

This technique is explained in the following video.

Plus forty/minus forty explained

This video derives the plus forty/minus forty technique

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Derivation of the plus forty/minus forty technique

Video 2.10.2: Graphing Part 2: temperature conversions

Example 2.10.7: Celsius to Fahrenheit

Convert 184K to Fahrenheit

Solution

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Note, this video needs to be redone when I get back to AR, as I forgot to write down the negative sign (although I stated it), and the date was wrong, as it was really 1983.

Ans: -128°F

Video 2.10.3: Temperature conversion Plus 40 minus 40 method (C to F)

Exercise 2.10.1

Add exercises text here.

Answer available at the end of this section.

Example 2.10.8: Fahrenheit to Celsius

136 °F =? °C

Solution

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

58°F

Video 2.10.4: Temperature conversion Plus 40 minus 40 method (F to C)

Answers

Exercise 2.10.1

We start by setting them equal to a common T,

$$T_F = T_C = T$$

Then substitute into the equation and solver for T

$$T = \frac{9}{5}T + 32$$
$$T(1 - \frac{9}{5}) = 32$$
$$T(\frac{-5}{5}) = 32$$
$$T = 32(\frac{-5}{4})$$
$$T = -40$$

.

Exercise 2.10.1

Add texts here. Do not delete this text first.

2.10.1: Graph Paper Page

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2.10.1: Graph Paper Page

Mass (g)	Volume (mL)
1.2g	0.7
17.8	10.0
29.0g	18.0
42.3	25.0
50.0g	30.0

Figure 2.10.1.1: 25x30 (hxw) blank sheet of graph paper for Ipad use in lecture

Mass (g)	Volume (mL)
1.2g	0.7
17.8	10.0
29.0g	18.0
42.3	25.0
50.0g	30.0

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3: ATOMS AND ELEMENTS

Introduction



- Topic hierarchy

3.1: The Nuclear Atom

3.2: The Elements

3.3: Isotopes

3.4: Isotopic Abundances

3.5: The Periodic Table

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3.1: The Nuclear Atom

The Nuclear Atom- Overview

The precise physical nature of atoms finally emerged from a series of elegant experiments carried out between 1895 and 1915. The most notable of these achievements was Ernest Rutherford's famous 1911 alpha-ray scattering experiment, which established that

- Almost all of the *mass* of an atom is contained within a tiny (and therefore extremely dense) ***nucleus*** which carries a positive electric charge whose value identifies each element and is known as the ***atomic number*** of the element.
- Almost all of the *volume* of an atom consists of empty space in which **electrons**, the fundamental carriers of negative electric charge, reside. The extremely small mass of the electron (1/1840 the mass of the hydrogen nucleus) causes it to behave as a quantum particle, which means that its location at any moment cannot be specified; the best we can do is describe its behavior in terms of the probability of its manifesting itself at any point in space. It is common (but somewhat misleading) to describe the volume of space in which the electrons of an atom have a significant probability of being found as the *electron cloud*. The latter has no definite outer boundary, so neither does the atom. The radius of an atom must be defined arbitrarily, such as the boundary in which the electron can be found with 95% probability. Atomic radii are typically 30-300 pm.

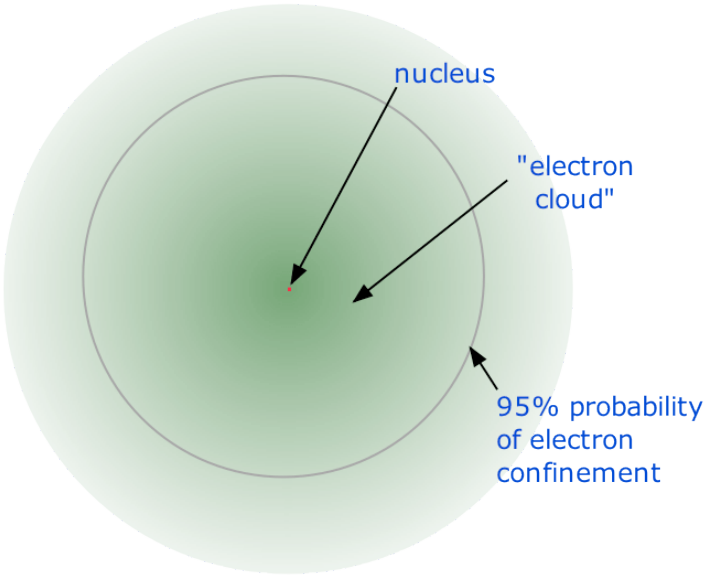


Figure 3.1.1: The structure of the nuclear atom with a central nucleus and surrounding electrons.

The nucleus is itself composed of two kinds of particles. ***Protons*** are the carriers of positive electric charge in the nucleus; the proton charge is exactly the same as the electron charge, but of opposite sign. This means that in any [electrically neutral] atom, the number of protons in the nucleus (often referred to as the *nuclear charge*) is balanced by the *same* number of electrons outside the nucleus. The other nuclear particle is the ***neutron***. As its name implies, this particle carries no electrical charge. Its mass is almost the same as that of the proton. Most nuclei contain roughly equal numbers of neutrons and protons, so we can say that these two particles together account for almost all the mass of the atom.

> Note
>
> Because the electrons of an atom are in contact with the outside world, it is possible for one or more electrons to be lost, or some new ones to be added. The resulting electrically-charged atom is called an *ion*.

Interactive Slide Show

The following slide show covers the history of the development of atomic theory from early Greek philosophers to modern atomic theory (click the arrow on the right). Prep Chem students will not be tested on this material, but they will need to know it in general chemistry, and interested students are advised to check out the general chemistry section 2.1.2 and 2.1.3.

Interactive Content

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View this Interactive Content online

Subatomic Particles

Atoms consist of electrons, protons, and neutrons. Although this is an oversimplification that ignores the other subatomic particles that have been discovered, it is sufficient for discussion of chemical principles. Some properties of these subatomic particles are summarized in Table 3.1.1, which illustrates three important points:

1. Electrons and protons have electrical charges that are identical in magnitude but opposite in sign. Relative charges of –1 and +1 are assigned to the electron and proton, respectively.
2. Neutrons have approximately the same mass as protons but no charge. They are electrically neutral.
3. The mass of a proton or a neutron is about 1836 times greater than the mass of an electron. Protons and neutrons constitute the bulk of the mass of atoms.

Table 3.1.1: Properties of Subatomic Particles (values from NIST Reference on Constants, Units, and Uncertainty).

Particle	Mass (g)*	Atomic Mass (amu)*	Electrical Charge (coulombs)	Relative Charge
electron	$9.1093837015\times10^{-28}$	0.0005485799090	$-1.602176634 \times 10^{-19}$	–1
proton	$1.67262192369 \times10^{-24}$	1.007276466621	$+1.602 \times 10^{-19}$	+1
neutron	$1.67492749804\times10^{-24}$	1.00866491595	0	0

Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

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- Modified by Joshua Halpern, Scott Sinex and Scott Johnson
- November Palmer & Ronia Kattoum(UALR)
- Elena Lisitsyna (H5P interactive modules)

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3.2: The Elements

The Elements

To date, 118 different elements have been discovered. Each element has a unique number of protons, which is the atomic number of an atom of that element. Students often get the terms atoms and elements mixed up. An atom is an entity of matter, it has a nucleus and electrons. An element is a type of atom, based on the number of protons in the atom's nucleus. Students in Prep Chem are required to memorize all the name and symbol of all the atoms that are not shaded in the periodic table of figure 3.2.1

UALR students in my class are required to know name and symbol of all the elements in this periodic table that are not shaded
If you click on the table it will open a printable PDF of the image.

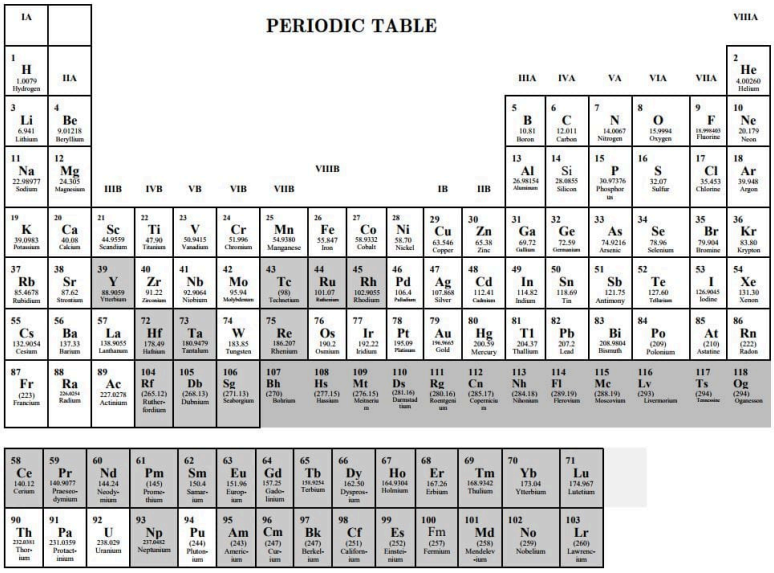


Figure 3.2.2: UALR students are required to know unshaded elements of this table, so start memorizing them early!

Click here for a Downloadable PDF

Naming Convention

In writing the symbol for an element you always capitalize the first letter and none of the others, and students will loss points if they do not follow this convention. For example, Co is different than CO, with the first being the element cobalt, and the second a molecule called carbon monoxide

Tough Ones

Many of the elements have their symbols derive from their common English names, like C for carbon, U for uranium and Ca for calcium. Elements have also been named for their properties [such as radium (Ra) for its radioactivity], for the native country of the scientist(s) who discovered them [polonium (Po) for Poland], for eminent scientists [curium (Cm) for the Curie], for gods and goddesses [selenium (Se) for the Greek goddess of the moon, Selene], and for other poetic or historical reasons. Some of the symbols used for elements that have been known since antiquity are derived from historical names that are no longer in use; only the symbols remain to indicate their origin. Examples are Fe for iron, from the Latin *ferrum*; Na for sodium, from the Latin *natrium*; and W for tungsten, from the German *wolfram*. Examples are in Table 2.1.2.

Table 2.1.2: Element Symbols Based on Names No Longer in Use

Element	Symbol	Derivation	Meaning
antimony	Sb	stibium	Latin for “mark”
copper	Cu	cuprum	from Cyprium, Latin name for the island of Cyprus, the major source of copper ore in the Roman Empire
gold	Au	aurum	Latin for “gold”
iron	Fe	ferrum	Latin for “iron”
lead	Pb	plumbum	Latin for “heavy”
mercury	Hg	hydrargyrum	Latin for “liquid silver”
potassium	K	kalium	from the Arabic al-qili, “alkali”
silver	Ag	argentum	Latin for “silver”
sodium	Na	natrium	Latin for “sodium”
tin	Sn	stannum	Latin for “tin”
tungsten	W	wolfram	German for “wolf stone” because it interfered with the smelting of tin and was thought to devour the tin

It is imperative that you memorize the names of the atoms of the elements that are not shaded in figure 3.2.1 and pay extra attention to the ones in table 3.2.1. On an exam you will probably not be asked what is the symbol of hydrogen or oxygen, but of radium or radon.

Test Yourself

Homework: Section 3.2

he following Flash Cards require you to type the names of 22 elements that are commonly misspelled.

Query 3.2.1

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For additional practice on names you can do the following quiz.

Query 3.2.1

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3.3: Isotopes

Atomic Number and Mass Number

When you study the periodic table, the first thing that you may notice is the number that lies above the symbol. This number is known as the **atomic number**, which identifies the number of protons in the nucleus of ALL atoms in a given element.

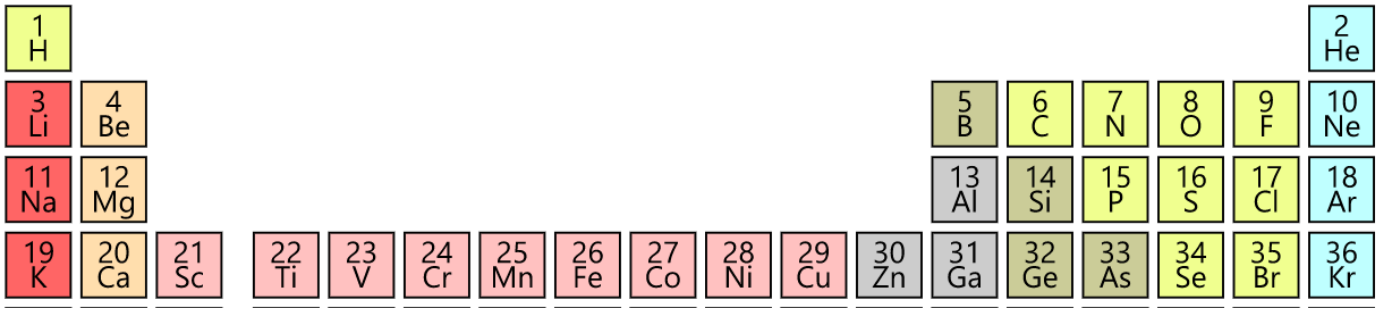


Figure 3.3.1: First four rows of the periodic table showing the atomic number Z for each of the elements, which is the number of protons an atom of that element contains.

The symbol for the atomic number is designated with the letter **Z**. For example, the atomic number (z) for sodium (Na) is 11. That means that all sodium atoms have 11 protons. If you change the atomic number to 12, you are no longer dealing with sodium atoms, but magnesium atoms. Hence, the atomic number defines the element in question.

Recall that the nuclei of most atoms contain neutrons as well as protons. Unlike protons, the number of neutrons is not absolutely fixed for most elements. Atoms that have the same number of protons, and hence the same atomic number, but different numbers of neutrons are called **isotopes**. All isotopes of an element have the same number of protons and electrons, which means they exhibit the same chemical behavior. Because different isotopes of the same element haves different number of neutrons, each of these isotopes will have a different **mass number(A)**, which is the sum of the number of protons and the number of neutrons in the nucleus of an atom.

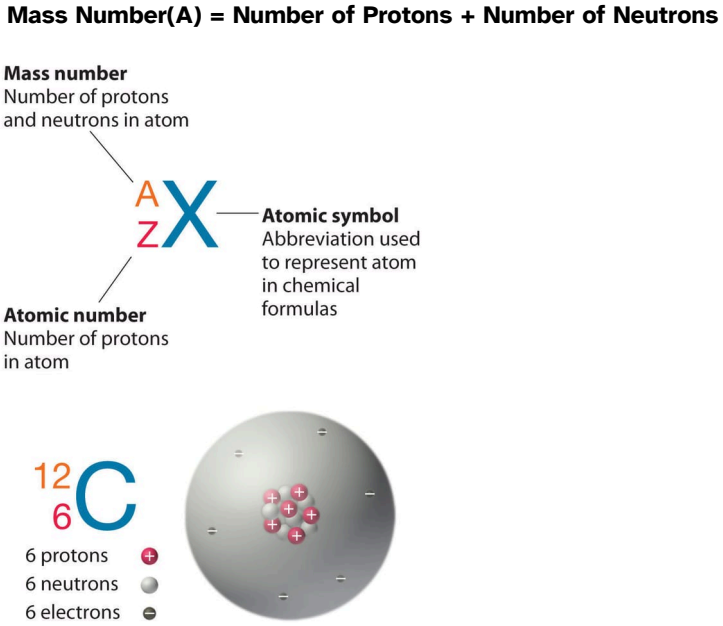


Figure 3.3.2: Formalism used for identifying specific nuclide (any particular kind of nucleus)

The element carbon (C) has an atomic number of 6, which means that all neutral carbon atoms contain 6 protons and 6 electrons.

In a typical sample of carbon-containing material, 98.89% of the carbon atoms also contain 6 neutrons, so each has a mass number of 12. An isotope of any element can be uniquely represented as $^{A}_{Z}X$, where X is the atomic symbol of the element, A is the mass number and Z is the atomic number. The isotope of carbon that has 6 neutrons is therefore $^{12}_{6}C$. The subscript indicating the atomic number is actually redundant because the atomic symbol already uniquely specifies Z. Consequently, it is more often written as ^{12}C, which is written as "**carbon-12,**" (^{14}C, would thus be written as "**carbon-14**.") Nevertheless, the value of Z is commonly included in the notation for nuclear reactions because these reactions involve changes in Z.

Note

The value of the atomic number Z is often ommited as the symbol defines it, so $^{12}_{6}C$. is often written as ^{12}C because any atom of the element carbon must have 6 protons, or it would not be carbon.

Atomic Mass Unit (Dalton)

The atomic mass unit (u, amu or Da) is essentially 1/12th the mass of a ^{12}C atom (that was the definition before 2019, and it is now based on the 7 immutable constants). This is a very small number and can be related to the macroscopic world through the following conversion factor.

$$1amu = 1.66054x10^{-24}g = 1.66054 \text{ yg} \tag{3.3.1}$$

(yg=yoctogram, see section 2.7- SI prefixes)

The amu is also called the Dalton and large molecules like proteins, nucleic acid and large polymers are often expressed in Kilo Daltons (1kDA=10^3 Da) or Mega Daltons (1MDa=10^6 Da).

Molar Mass

The atomic mass unit is very small, as it is effectively a scale based on the mass of an atom, and in the world we live and breath in measurements are of the order of magnitude of a grams and kilograms, and so taking the reciprocal of 3.31 we can get a feel for how many atomic mass units are in a gram.

$$1g = \frac{amu}{1.6605x10^{-24}} = 6.0221x10^{23}amu \tag{3.3.2}$$

The value of 6.022x10^{23} is also known as Avogadro's constant, as representing the number of entities in the SI base unit of the mole (section 2.4.3)

$$1\text{mole of anything} = 6.022x10^{23}\text{things} \tag{3.3.3}$$

The Molar Mass of a substance is the mass of one mole of that substance (6.02x10^{23} particles) and so the atomic weight in AMUs is to an atom what the molar mass is in grams to a mole of atoms. That is, one atom of ^{1}H weighs 1 amu, while one mole of ^{1}H weights 1g.

$$\begin{aligned} 1\text{atom H} &= 1amu \\ 1\text{mole H} &= 1g \end{aligned} \tag{3.3.4}$$

Isotopes

Isotopes are different atoms of the same element that contain different numbers of neutrons in their nucleus, and therefore they have different atomic masses. For a neutral atom the number of electrons = the number of protons.

$$\begin{gathered} {}^{54}_{26}\mathrm{Fe} \text{ which is often written as } {}^{54}\mathrm{Fe} \\ \text{has} \\ 26 \text{ protons, } 28 \text{ neutrons and } 22 \text{ electrons} \end{gathered} \tag{3.3.5}$$

For a charged atom the number of protons is not equal to the number of electrons and you have an ion. There are two types of ions, **cations** [+] charged and **anions** [-] charged. We write the charge as a post superscript. The charge of an ion is the number of protons minus the number of electrons, as each proton has a charge of [+1] and each electron has a charge of [-1].

Cations

Number of electrons is less than the number of protons

$$\begin{gathered} {}^{54}_{26}\mathrm{Fe}^{+2} \text{ which is often written as } {}^{54}\mathrm{Fe}^{+2} \\ \text{has} \\ 26 \text{ protons, } 28 \text{ neutrons and } 24 \text{ electrons} \end{gathered} \tag{3.3.6}$$

Anions

Number of electrons is greater than the number of protons.

$$\begin{gathered} {}^{37}_{17}\mathrm{Cl}^{-} \text{ which is often written as } {}^{37}\mathrm{Cl}^{-} \\ \text{has} \\ 17 \text{ protons, } 20 \text{ neutrons and } 18 \text{ electrons} \end{gathered} \tag{3.3.7}$$

Example 2.2.1

An element with three stable isotopes has 82 protons. The separate isotopes contain 124, 125, and 126 neutrons. Identify the element and write symbols for the isotopes.

Given: number of protons and neutrons

Asked for: element and atomic symbol

Strategy:

1. Refer to the periodic table and use the number of protons to identify the element.
2. Calculate the mass number of each isotope by adding together the numbers of protons and neutrons.
3. Give the symbol of each isotope with the mass number as the superscript and the number of protons as the subscript, both written to the left of the symbol of the element.

Solution:

A The element with 82 protons (atomic number of 82) is lead: Pb.

B For the first isotope, A = 82 protons + 124 neutrons = 206. Similarly, A = 82 + 125 = 207 and A = 82 + 126 = 208 for the second and third isotopes, respectively. The symbols for these isotopes are ${}^{206}_{82}\mathrm{Pb}$, ${}^{207}_{82}\mathrm{Pb}$, and ${}^{208}_{82}\mathrm{Pb}$, which also can also be symbolized as Pb-206, Pb-207, and Pb-208.

Exercise 3.3.1

What is the mass number of a phosphorous atom with 16 neutrons

Answer available at the end of this section.

Exercise 3.3.2

How many protons, neutrons, and electrons are in As-74 atom?

Answer available at the end of this section.

Exercise 3.3.3

The actual mass of As-74 is 73.924 amu's. (1 amu= 1.66054 x 10^{-24} g)

A) What is the mass of As-74 in grams?

B) What is the mass of the arsenic atom relative to the carbon-12 atom?

Answer available at the end of this section.

Test Yourself

Homework: Section 3.3

Graded Assignment: Section 3.3

Query 3.3.1

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Contributors and Attributions

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- November Palmer
- Elena Lisitsyna (H5P interactive modules)

Answers

31

33 protons, 41 neutrons, 33 electrons

A) Mass of As-74 = (73.924 amu) x (1.66054 x 10^{-24} g/amu) = 1.2275 x 10^{-22} g

B) Mass of As-74/Mass of C-12 = 73.924 amu/12.000 amu= 6.1603 ... so, the mass of an As-74 atom is 6.1603 times heavier than a C-12 atom.

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3.4: Isotopic Abundances

Learning Objectives

Goals:

- Understand Isotopic Distribution of the atoms for various elements as they exist in their natural state
- Recognize the meaning of the atomic mass on the periodic table

By the end of this module, students should be able to:

- Determine the average atomic mass from the natural isotopic distribution of the atoms of an element
- Using the atomic mass on the periodic table for an element with two isotopes, and knowing two of the four data values of % composition and isotopic mass of each isotope, determine the other two values.

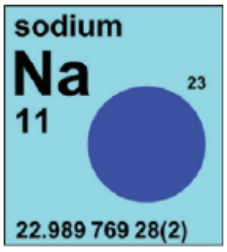
Prior knowledge:

- Section 2.9: Fractions and Percent
- Section 3.3: Isotopes

Atomic Weight

The atomic weight is the mass of an atom, typically expressed in atomic mass units (amu). For an isotope, it is the mass of the nucleus, that is the mass of the protons and neutrons, as the mass of the electrons are considered negligible. The atomic weight is often posted on the periodic table under the elment's symbol. In their natural state only 21 elements exist as single isotopes (see section 2.2.2 of Gen Chem 1), and most elements exist as a mixture of nuclei from multiple isotopes.

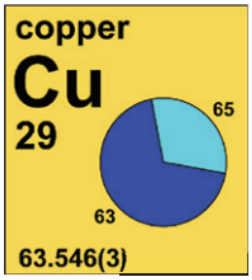
Sodium is one of the elements that exists as a single isotope, ^{23}Na.



Stable Isotope	Relative Atomic Mass	Mole Fraction
^{23}Na	22.989 769 29	1

Figure 3.4.1*:* Sodium only has one isotope, and note, the uncertainty of this value is expressed by the (2) in 22.989 769 28(2), indicating that the last value has an uncertainty of +/- .000 000 02. Adapted from 2018 IUPAC Technical Report - IPTEI for Education Community, (cc. 4.0), p. 1864.

For elements that have more than one isotope the atomic weight is the average weight based on the fractional abundance of each isotope, and this is the value given on the periodic table. Copper has two isotopes, ^{63}Cu (69.15%, mass=62.9300 amu) and ^{65}Cu (30.85%, mass = 64.928 amu), and so the respective mole fractions are 0.6915 and 0.3085, resulting in an average atomic weight of 63.55 amu, even though there is not a single atom that weighs 63.55 amu.



$$\underbrace{\frac{0.6915}{fraction\ ^{63}Cu}}\ \underbrace{(62.9300\,amu)}_{mass\ ^{63}Cu} + \underbrace{\frac{0.3085}{fraction\ ^{65}Cu}}\ \underbrace{(64.928\,amu)}_{mass\ ^{65}Cu} = \underbrace{63.55\,amu}_{\text{average mass}}$$
$$note:\ 0.6915 + 0.3085 = 1$$

Figure 3.4.2: Natural samples of copper contain two isotopes, and its atomic weight is to four significant digits is 63.55 amu, even though there is not a single atom of copper that weights 63.55 amu.

Exercise 3.4.1

True or False: The atomic weight of copper is 63.456 amu, which means an atom of copper weights 63.546 amu.

Answer available at the end of this section.

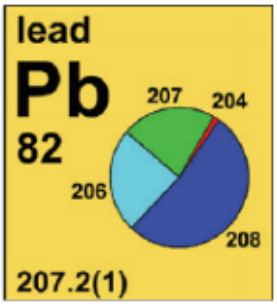
Exercise 3.4.2

The atomic weight of chorine is _____and the atomic number of chlorine-35 is_____.

a. 35, 17
b. 17, 35
c. 35.4527; 17
d. 35.4527; 35

Answer available at the end of this section.

Lead has 4 stable isotopes as shown in Figure 3.4.3 . Lead is a fairly large element and has many radioisotopes, in fact it has 29 isotopes with a half life of less than one hour, seven between an hour and a year, and two with half lives over a year (^{202}Pb = 5.25x10^{4} years & ^{210}Pb = 22.3 years), but these are insignificant to the thousandth position (note the four mole fractions add to one).



Stable Isotope	Relative Atomic Mass	Mole Fraction
^{204}Pb	203.973 043	0.014
^{206}Pb	205.974 465	0.241
^{207}Pb	206.975 897	0.221
^{208}Pb	207.976 652	0.524

Figure 3.4.3: Lead has 4 stable isotopes. Adapted from 2018 IUPAC Technical Report - IPTEI for Education Community, (cc. 4.0), p 2007.

Exercise 3.4.3

Nitrogen has two isotopes, the first isotope has a mass of 14.0004amu, while the second isotope has a mass of 15.000 amu, which isotope has the greatest natural abundance? Tip, you need to look at the periodic table.

a. ^{14}N
b. ^{15}N
c. There are equal amounts.
d. Not enough information provided.
e. none of the above

Answer available at the end of this section.

Isotopic Abundance Calculations

For elements with more than one isotope there are a variety of problems you need to be able to solve. You not only need to be able to calculate the average mass from the isotopic abundance and masses, but go backwards, using the atomic weight on the periodic table as the average atomic mass. Rewriting eq. 3.4.1 in terms of algebraic variables gives us a feel for the types of problems (review section 2.9 Fractions and Percent). Consider an element with two isotopes A & B, each having a mass of m_A or m_b and a mole fraction of X_A or X_b.

$$\underbrace{X_A}_{\text{fraction A}}\ \underbrace{(m_A)}_{\text{mass A}} + \underbrace{X_B}_{\text{fraction B}}\ \underbrace{(m_B)}_{\text{mass B}} = \underbrace{m_{ave}}_{\text{average mass}} \tag{3.4.2}$$

In general the equation for an element with [n] isotopes is:

$$\sum_{i=1}^{n} X_i m_i = m_{ave} \tag{3.4.3}$$

Note, the equation for two isotopes has 5 unknowns, and you can get m_{ave} from the periodic table, so you need to be given three of the other variables to solve the equation (one equation can only be solved for one unknown). But there is another equation we can use to reduce the number of unknowns. That is the sum of the fractions equals 1

$$X_A + X_B = 1 \tag{3.4.4}$$

or in general

$$\sum_{i=1}^{n} X_i = 1 \tag{3.4.5}$$

(this is the same equation as 2.9.4 when we looked at the fraction of salt in salt water).

Example 3.4.1: Calculating Average Atomic Mass

What is the average atomic mass of Neon, given that it has 3 isotopes with the follow percent abundances;

^{20}Ne = 19.992 amu (90.51%), ^{21}Ne = 20.993 amu (0.27%), ^{22}Ne = 21.991 amu.

What we know: since you know what the element is, you can solve this without doing any math by using the periodic table, but you need to be able to do the math because it might be an unknown, and that is the only way you can Figure out the correct significant Figures.

Since Ne-20 has the greatest percent abundance, it should have the most impact on your average. Therefore, we expect the average atomic mass to be closer to the mass of Ne-20 (about 19.992 amu). Click the following video tutor to see if we estimated correctly.

YouTube Video

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Video 3.4.2: 4:53 minute YouTube solving above problem.

Answer According to the correct number of significant Figures, we came up with 20.18 amu as the average atomic weight even though the average atomic weight from the periodic table is 20.179 amu. However, it is still a good check to make sure that you are on the right path.

Check Yourself: We predicted earlier that our answer should be closer to the mass of Ne-20 (19.992 amu) instead of Ne-21 or Ne-22 because it has the greatest natural abundance, and thus, impacts the average more. We can see that the math does align with our logic!

The following problem is more complicated. This is mathematically the same problem as equation \({3{.}4{.}\eatSpaces{2}\endSpaces}) or 2.9.4 of the last chapter, and has 5 variables, but you have only been given two (fraction of ^{35}Cl and the mass of ^{35}Cl) and you are asked what is the mass of the other isotope. So you can algebraically solve it for the other isotope (M_B of equation 3.4.2), but your algebraic solution would have two variables in it (fraction or percent of the other isotope and the average mass). By now you should realize the value of the average is what is given on the periodic table, and so you can look that up, and you can solve for the fraction of the second isotope by noting the sum of the fractions equals 1 ($eq\(3.4.4$). Try to follow this problem on your own, and then watch the video

Example 3.4.2: Isotopic Mass Calculation

Chlorine has two isotopes, with 75.53% being ^{35}Cl with an isotopic mass of 34.969 amu, what is the mass of the other isotope?

What we know: In this case, you have the average atomic mass (from the periodic table). You are trying to find the mass of the individual isotope. You also know that the individual isotopes have to add up to 100%.

YouTube Video

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Answer: The correct answer is 36.9 amu.

Test Yourself

Homework: Section 3.4

Graded Assignment: Section 3.4

Query 3.4.1

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Answers

Exercise 3.4.1

False, 63.456 amu is the average atomic weight of copper, there are actually two isotopes, one weighing 62.930 amu and one weighing 64.928 amu.

C) the atomic weight is the average of mass of all isotopes of chlorine atoms and found below the symbol on the periodic table. The atomic number is the number of protons in all chlorine atoms and is found on the top of the symbol in the periodic table.

A) The average atomic weight is 14.007 amu which is much closer to 14 then 15, so 14 is the most abundant isotope. In the next section you will use this kind of data to calculate fractional abundances.

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3.5: The Periodic Table

Learning objectives

1. Identify the groups of the periodic table
2. Identify elements as metals, nonmetals, metalloids
3. Infer likely properties of elements based on their position on the periodic table.

Classification of the Periodic Table

If you tried to click on any of the above elements, the details of each element may seem overwhelming and it may seem incredibly difficult to find patterns. However, the beauty of the periodic table is found in it's *periodicity,* repeating patterns that help us predict physical and chemical properties of various elements. Thus the periodic table can also be split into seven rows, known as **periods** (Figure 3.5.4). At first glance, it may seem like the two lone rows at the bottom of the period table, known as the **actinide and lanthanide series** should be periods 8 and 9. However, if you look carefully, these two row should actually be wedged in the 6th and 7th period, respectfully, in order to follow the order of increasing atomic number. Those rows are kept at the bottom to decrease the width of the periodic table and are also known as the **inner transition metals.**

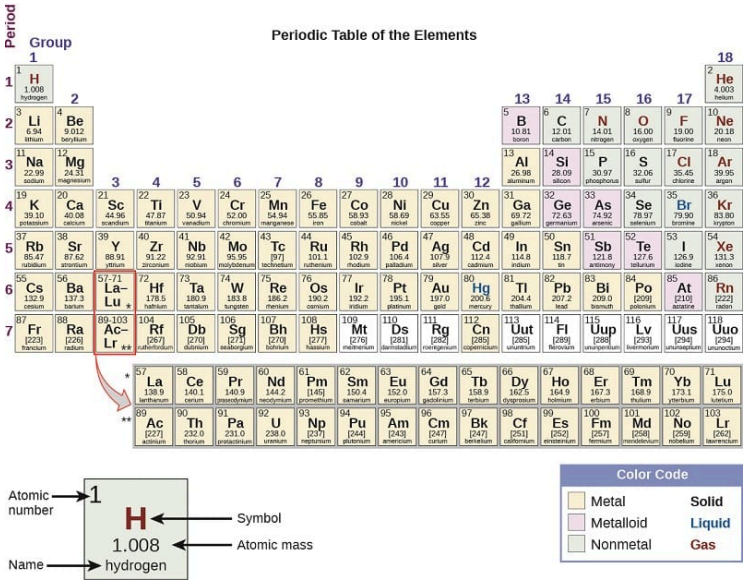


Figure 3.5.1: The periodic table is split into seven periods(rows) with the **inner transition metals** shown separately on the bottom to decrease the width of the periodic table. It can also be broadly divided into metals, nonmetals, and metalloids(semimetals). The states of the elements is also indicated at room temperature of 25^{o} C.

Groups can be numbered as 1-18 (Figures 3.5.1 or 3.5.2) or they can be numbered 1-8 followed by "A" or "B" (Figure 3.5.3). The "A" designation represents the **main group elements** (Columns 1-2, and Columns 13-18) The "B" designation represents **transition elements** (Columns 3-12). Either group numbering system is acceptable, but we will see later that there is a benefit for knowing the group numbers with the "A" and "B" designation. In addition, some of the individual groups have a particular group name. For example, Group 17 or 7A elements are also known as the **halogens.** Group 2 or 2A elements are known as the **alkaline group metals.** Group names are commonly used by chemists and are worth knowing. But, most importantly, elements in the same groups tend to have similar chemical and physical properties. We will also discuss the reason behind such patterns later in this text. Figure 3.5.2 provides the common names of the groups.

Exercise 3.5.1

Which of the following elements will most likely have similar chemical properties to fluorine?

A) neon

B) nitrogen

C) iodine

D) sulfur

Answer available at the end of this section.

The Families of the Periodic table

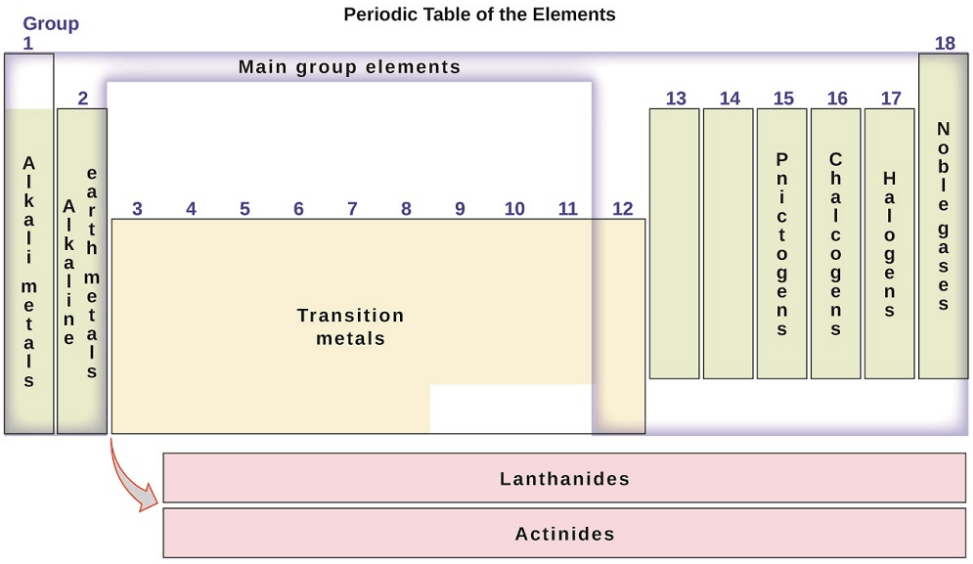


Figure 3.5.2: Families or groups of the Periodic Table

Exercise 3.5.2

What alkali metal is found in period 4?

A) carbon

B) potassium

C) calcium

D) Titanium

Answer available at the end of this section.

Figure 3.5.3: Modern Family or group numbering system for the Periodic Table along with the names of the families.

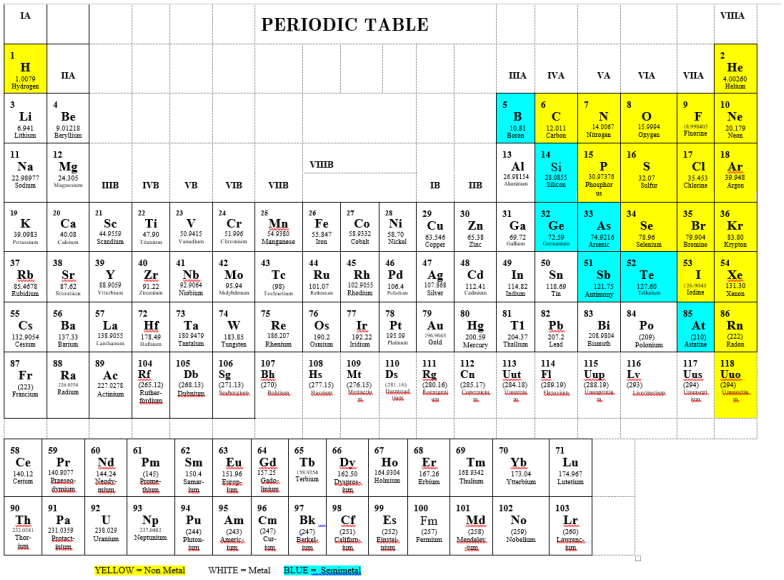


Figure 3.5.3: The older method of numbering the families or groups. Note the color coding of this table indicates metals, metaloids (semi-metals) and non-metals. These will be discussed below.

Exercise 3.5.3

What is the group number for oxygen?

A) 16

B) 6A

C) 16A

D) Both A and B

Answer available at the end of this section.

If you start at boron and draw "stairs" down to astatine, the seven elements that are "on" are the **metalloids or semi-metals** (Figures 3.5.3). Everything to the left of the "stairs" with the exception of hydrogen is a **metal**. Everything to the right of the "stairs", in addition to hydrogen, is a **nonmetal.** This is essential for describing common physical properties for each of the three categories and will be used in nomenclature of chemical compounds.

Exercise 3.5.4

What is true regarding the classification of the following silicon? (select all that apply)

A) group 4A

B) metalloid

C) main group element

D) period 14

E) nonmetal

Answer available at the end of this section.

Physical Properties of Metals:

Note, these can be related to the fact that electrons in metallic bonds are loosely bonded between multiple core atoms and freely flow from one atom to another.

1. Conduction (heat and electricity)
2. Malleability (can be hammered into thin sheets)
3. Ductility (can be pulled into wires)
4. Lustrous (Shiny appearance)

Physical Properties of Non-Metals:

Note, nonmetals are held together by directional covalent bonds of shared electrons, and this relates to their properties.

1. Poor Conductors
2. Nonlusterous
3. Nonmalleable
4. Brittle
5. Brightly colored

They have both properties of metals and non-metals. One of the most important physical properties of metalloids is their semi-conductive properties.

Natural States of Atoms

If you go to the Pubchem periodic table and click "standard state, you will see 2 elements are liquid, 11 are gasses, and one (Og) is "expected" to be a gas. If you click on Og and go to it's PubChem elemental page you see the most stable isotope has a half life of 0.89 milliseconds, and only a few atoms have ever been produced. You are required to know these, and note, several elements like Ga, Cs and Fr melt at very low temperatures and so may actually be liquids at room temperature

Liquids: **Hg, Br_2,** (Ga, mp=85.58^{o}F; Cs, mp=83.3^{o}F; Fr, mp=80.6^{o}F and Rb, mp=103.1 ^{o}F)

Gasses: Noble Gasses and Lighter diatomics, **H_2, N_2, O_2, F_2, & Cl_2**

Solids: All other elements

Vocabulary

Test Yourself

Homework: Section 3.5

Query 3.5.1

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- Anonymous
- Modifications by Ronia Kattoum
- Elena Lisitsyna (H5P interactive modules)

Answers

C) iodine. Elements in the same **group** have similar chemical properties, not the same periods.

B) Potassium is in Group (Column) 1A and is in the 4th period (row).

D

A, B, and C are all correct. D is incorrect. The period (row) is 3. The group is 13. E is incorrect because silicon sits on the stairs and is one of the metalloids.

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4: COMPOUNDS AND MOLECULES

4.1: Chemical Compounds
4.2: Molecular Compounds
4.3: Ionic Compounds and Formulas
4.4: Polyatomic Ions
4.5: Nomenclature

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4.1: Chemical Compounds

learning objectives

1. Differentiate between ionic compounds and molecular compounds in the way the bonds form and varying characteristics
2. Identify elements that exist as diatomic molecules
3. Define allotropes
4. Interpret different types of formulas including molecular formulas, structural formulas, and condensed structural formulas

Introduction

Chemical compounds occur when 2 or more atoms are bonded together. We will consider two basic types of bonds, covalent and ionic, and thus focus on two basic kinds of chemical compounds, covalent and ionic. Each of these compounds is associated with a chemical formula, and it is important to note that there is a third type of bond that is very common and which we are not considering here, and that is the metallic bond. A pure metal, or an alloy like Brass, (a mixture of copper and zinc) is neither an ionic or covalent compound. Brass does not form a "compound" in the classical sense of Dalton's theory, where there is a constant ratio of copper to zinc that leads to a specific formula. In fact brass can be considered to be a solid solution, and the different grades of brass depend on the ratio of copper to zinc. So we do not associate a typical chemical formula with metals and alloys.

Chemical bonds are the result of Coulombic (electrostatic) interactions between charged particles, where like charges repel and opposite charges attract. In the case of covalent bonds, these are between electrons and the nuclei of one atom with the electrons and nuclei of another. In the case of ionic bonds, these are between cations ([+] charged ions) and anions ([-] charged.

Types of Bonds

There are actually 4 types of bonds that commonly occur, covalent, ionic, polar covalent and metallic. At this point in our study we will focus on two types, covalent and ionic. The type of bond that forms between two atoms depends on the type of atoms and can often be determined by identifying if the bonding atoms are metallic or covalent. So we should start with figure 3.5.4 of the last chapter, which we are reposting below as figure 4.1.1.

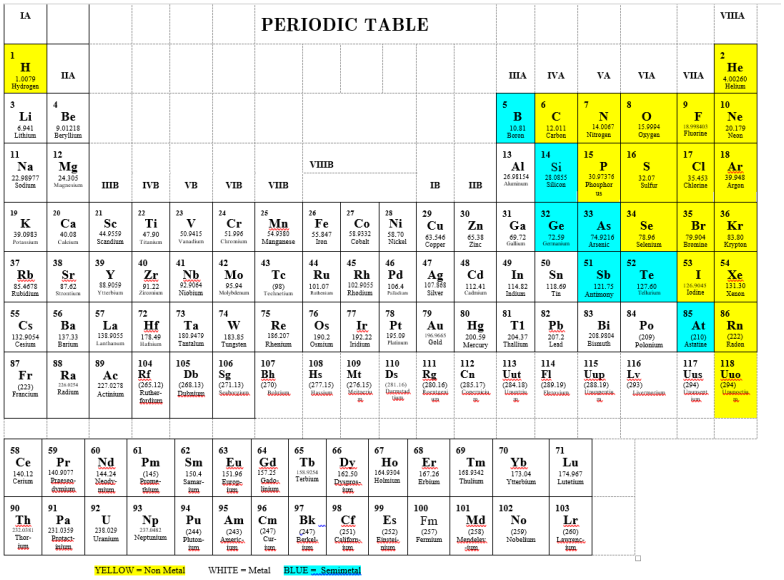


Figure 4.1.1: Periodic table distinguishing metals from nonmetals. Note this table uses the "old" way of labeling the families where it breaks them into two groups, the "A" group and the "B" group.

Covalent Molecules

Covalent Bonds occur between two nonmetals and involving a sharing of valence (outer shell) electrons. These are the classical molecules and can range from simple molecules like water to very complex molecules like proteins. We will cover these in section 2 of this chapter. You will be responsible for naming simple covalent compounds (last section of this chapter). The molecular formula is the formula of the discrete entity of matter that a covalent compound represents. So H_2O represents a water molecule where the entity of matter is two hydrogens bonded to one oxygen.

Ionic Compounds

Ionic compounds occur between a metal and a nonmetal, and involving the formation of ions. The metal gives an electron to the nonmetal, and the positive charge metal is called a cation and the negative charged nonmetal is called an anion

- **Cation: Net positive charge,** the number of protons is greater than the number of electrons
- **Anion: Net negative charge,** the number of protons is less than the number of electrons.

Although true, the above statement that ionic compounds form between metals and nonmetals is a bit simplistic, and not all ionic compounds a formed between metals and nonmetals, but it is a good rue of thumb if you have a compound between two different elements. As we will see later in this chapter, some ionic bonds involve polyatomic ions, which are covalently bonded molecules that have a net charge.

Ionic compounds form crystal lattice structures in three dimensional space where the geometry is a result of the attraction between the opposite charged ions and the repulsion between the like charged ions, and this structure can extend in space until all the available ions are used up. If you look at a salt crystal you will see that each positive ion is surrounded on all 6 sides by negative ions, and that a crystal lattice forms. So it makes no sense to have a "molecular formula" for an ionic compound, because the number of atoms will grow as the crystal grows.

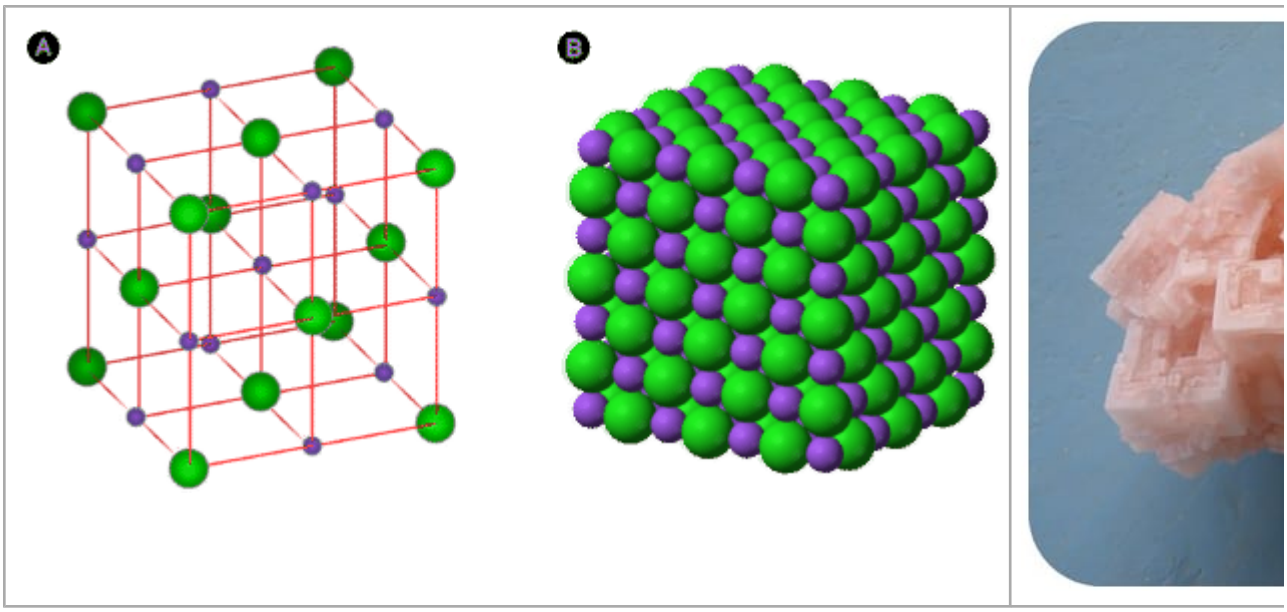


Figure 4.1.2 On the left is a crystal lattice of NaCl table salt, which is also shown on the right, where it forms Halite crystals. Because the number of ions depends on the size of the crystal, we use the lowest ratio of positive to negative ions that results in a neutral structure.

Principle of Charge Neutrality

This is a very important concept, which states that the formula of an ionic compound must be neutral as there were a net charge, the repulsive forces between the like charged ions would be greater in number than the attractive forces between opposite charged ions and so the material woud never form a stable structure. If we look at figure 4.1.2 we see that each sodium ion is within the bulk of the cube is bonded to 6 chloride ions, while each chloride ion is bonded to each sodium, and the total number of sodium and chloride ions is a consequence of the size of the crystal and so there is no discrete formula that describes sodium chloride the way a molecular formula describes a molecule like water. As a consequence we use the lowest whole number ratio of cations ([+] ions) to anions ([-] ions) to describe an ionic compound and the formula of sodium chloride is NaCl and not Na_6Cl_6. That is, the formula is the lowest whole number ratio of cations to anions that results in a neutral structure, and not a representation of the actual strucute itself.

Polar Covalent Compounds

Polar covalent compounds are covalent compounds with unequal sharing of the bonding electrons. They can behave very differently than pure covalent compounds that have equal sharing of the bonding electrons. Acids are a type of polar covalent compound and the following video shows the difference between covalent, ionic and polar covalent compounds. Please note the video goes over some concepts we will study later (electronegativity, and polar bonds), but if you watch between eleven seconds and 49 seconds you get a good visualization of the difference between a ionic and covalent bond.

YouTube Video

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View this YouTube Video online

Video 4.1.3: 2:11 YouTube describing polar bonds from the perspective of electronegativity. We will learn about electronegativity and polar bonds later.

Metallic Bonds

Metallic bonds occur between metals. If the metals are different, they are called alloys. Bonze is an alloy of copper and tin, while brass is an alloy of copper and zinc, and the properties of the alloys can vary as the ratio of the different metals varies. What is important to know is that metals form two types of bonds, metallic bonds between metals and ionic bonds with nonmetals. The iron of a steel fence is a metallic bond, and the rust on the iron, which forms when it reacts with the nonmetal oxygen, is an ionic bond. It should be noted that rust is very complex and there is no single formula that describes it

Test Yourself

Homework: Section 4.1

Query 4.1.1

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4.2: Molecular Compounds

Covalent Bonds and Molecules

Covalent bonds occur when two or more atoms *share* valence (outer shell) electrons resulting in the formation of **molecular compounds**, which typically involve nonmetals. These come about as a result of the Coulombic forces of attraction and repulsion between the [+] nuclei and the [-] valence (outer shell) electrons of different atoms. It is easiest to see this in the simplest molecule, the diatomic hydrogen. On the left of the diagram below, you have two isolated atoms that are so far away that the blue electron does not feel an attraction to the red nucleus, and the red electron does not feel an attraction to the blue nucleus, and so no bond forms. As they get closer together the electrons of each atom begin to feel an attraction to the nucleus of the other atom and they move closer together.

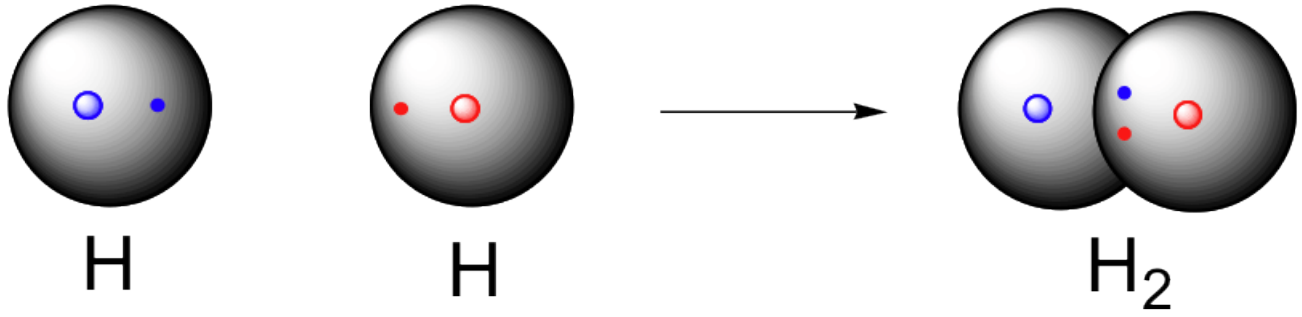


Figure 4.2.1: Simplified picture showing hydrogen as two isolated atoms (left) and as a diatomic covalent molecule (right)

Diatomics

Atoms of some elements form molecules by forming bonds with themselves, and if there are two atoms, they are called the diatomic elements, and you need to know the seven diatomics.

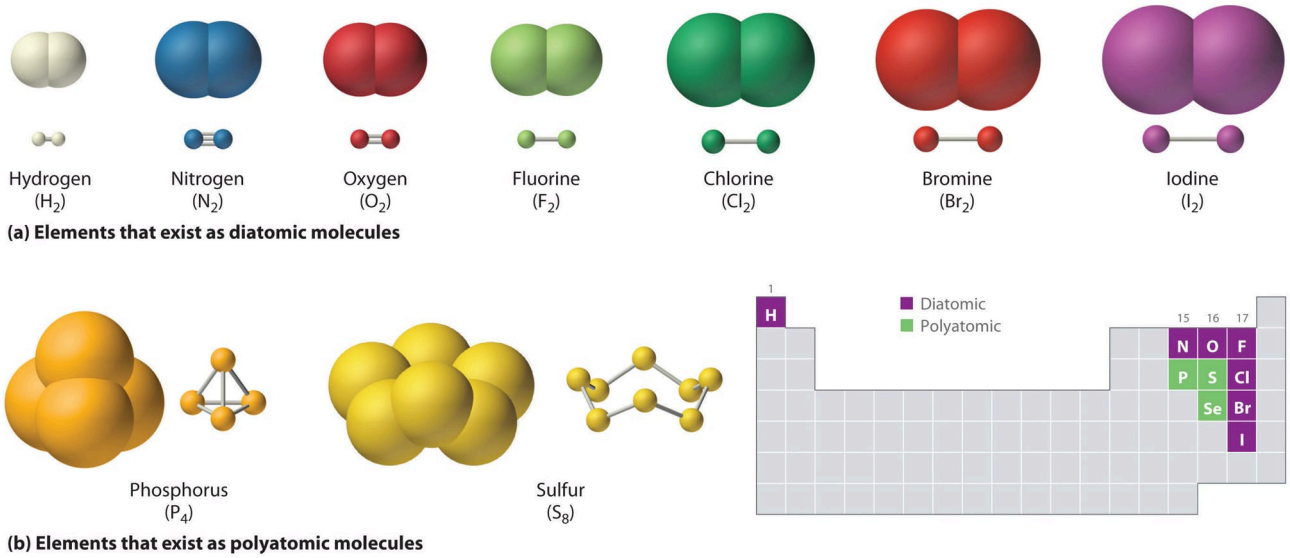


Figure 4.2.2: Elements that form diatomic and polyatomic molecules in their natural state (CC BY-SA-NC; Anonymous).

The ones on the periodic table that are purple form diatomic molecules, the ones in green form polyatomic molecules.

NOTE: When we say "oxygen" (or any diatomic molecule), we may mean monatomic oxygen O, or diatomic oxygen, O_2. We have to

look at the context of the statement. If we are describing it as a substance, like "he breathes oxygen", we mean O_2, but if we are describing it as part of a compound, like how many atoms of oxygen are in 3.00 grams of water, we mean O.

Allotropes

Allotropes are when atoms of an element can form more than one stable compound. For example, oxygen can form O_2 (diatomic oxygen) and O_3 (ozone), which are allotropes. (Not to be confused with isotopes, which are different nuclides of the same element containing different numbers of neutrons). Selenium forms several allotropes including Se_2, Se_6 and Se_8. Note, the most stable allotrope can change as you change the temperature and pressure, for example, at high T (900 ^{o}C) Se_2 is the most stable, but at lower temperature the larger ones became more stable.

YouTube Video

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Video 4.2.2: *1:51 min YouTube on Allotropes of Carbon uploaded by Elearnin (https://youtu.be/sUqr6Uk29Z8).*

Molecular Formula

When multiple atoms are held together by covalent bonds they form a single chemical entity, which we call a molecule. As postulated in Dalton's atomic theory, the ratio of the atoms of the different elements are whole numbers, and this can be described by the molecule's molecular formula. So each water molecule (H_2O) has two atoms of hydrogen and one of oxygen, while each molecule of aspirin ($C_9H_8.O_4$) has 9 carbon, 8 hydrogen and 4 oxygen.

Representations of Molecular Structures

Molecular formulas tell you how many atoms of each type of element are in a bond, but they do not necessarily tell you which atoms are bonded to which other atoms, or their orientation in space. In contrast, **structural formulas** show which atoms are bonded to one another and, in some cases, the approximate arrangement of the atoms in space. Knowing the structural formula of a compound enables chemists to create a three-dimensional model, which provides information about how that compound will behave physically and chemically.

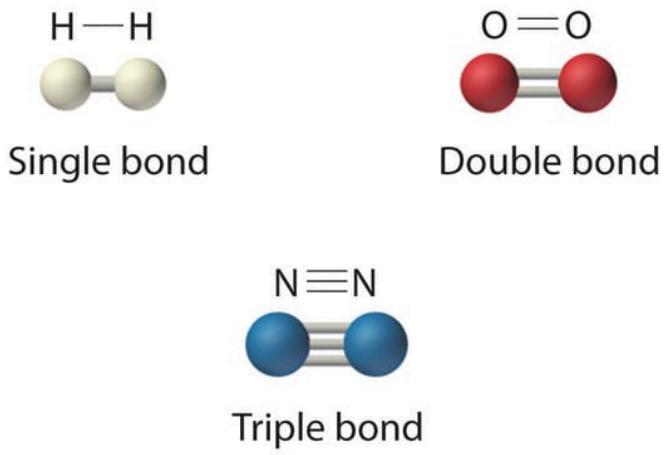


Figure 4.2.3: Molecules That Contain Single, Double, and Triple Bonds. Hydrogen (H_2) has a single bond between atoms. Oxygen (O_2) has a double bond between atoms, indicated by two lines (=). Nitrogen (N_2) has a triple bond between atoms, indicated by three lines (≡). Each bond represents an electron pair. *(image from LibreText core content.)*

The structural formula for H_2 can be drawn as H–H and that for I_2 as I–I, where the line indicates a single pair of shared electrons, a single bond. Two pairs of electrons are shared in a double bond, which is indicated by two lines, for example O_2 is O=O. Three electron pairs are shared in a triple bond, which is indicated by three lines, for example, N_2 is N≡N (see Figure 4.2.2 pr 4.2.3). Carbon is unique in the extent to which it forms single, double, and triple bonds to itself and other elements. The number of bonds formed by an atom in its covalent compounds is not arbitrary. Hydrogen, oxygen, nitrogen, and carbon have very strong tendencies to form substances in which they have one, two, three, and four bonds to other atoms, respectively (Table 4.2.1).

Table 4.2.1: The Number of Bonds That Selected Atoms Commonly Form to Other Atoms

Atom	Number of Bonds
H (group 1)	1
O (group 16)	2
N (group 15)	3
C (group 14)	4

The structural formula for water can be drawn as follows:

d

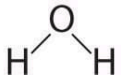


Figure 4.2.4: Structure of water

Because the latter approximates the experimentally determined shape of the water molecule, it is more informative. Similarly, ammonia (NH_3) and methane (CH_4) are often written as planar molecules, even though they are not planar:

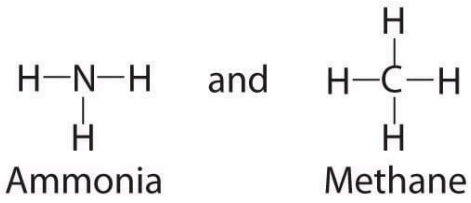


Figure 4.2.5: Structure of Ammonia and Methane

As shown in Figure 4.2.5, however, the actual three-dimensional structure of NH_3 looks like a pyramid with a triangular base of three hydrogen atoms. The structure of CH_4, with four hydrogen atoms arranged around a central carbon atom as shown in Figure 4.2.3, is tetrahedral: the hydrogen atoms are positioned at every other vertex of a cube. Many compounds—carbon compounds, in particular—have four bonded atoms arranged around a central atom to form a tetrahedron.

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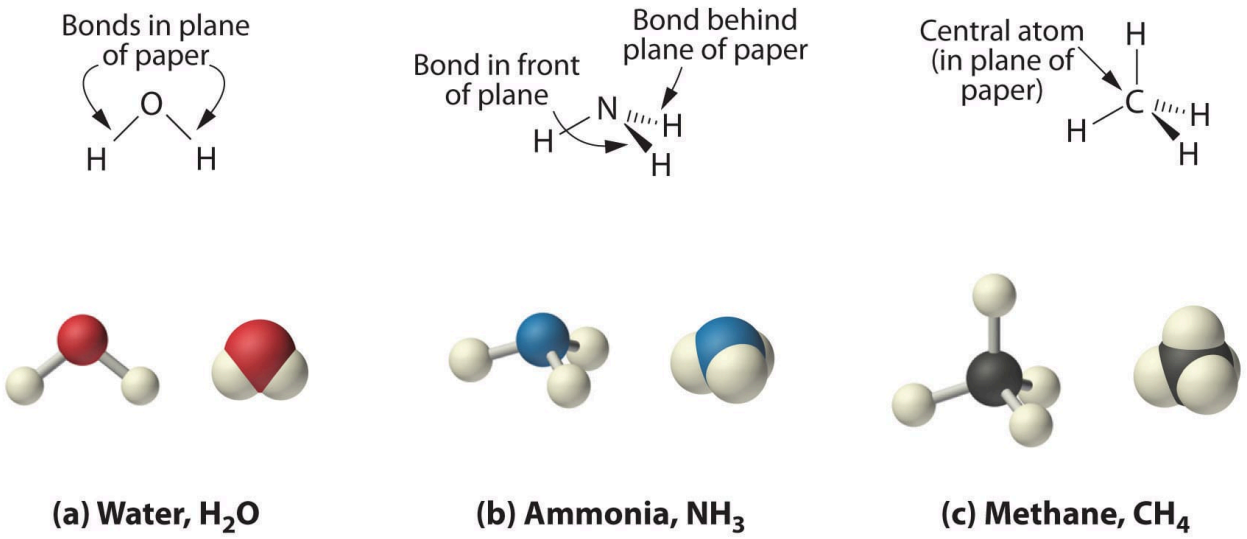


Figure 4.2.6:

Structure of Ammonia and Methane

The Three-Dimensional Structures of Water, Ammonia, and Methane. (a) Water is a V-shaped molecule, in which all three atoms lie in a plane. (b) In contrast, ammonia has a pyramidal structure, in which the three hydrogen atoms form the base of the pyramid and the nitrogen atom is at the vertex. (c) The four hydrogen atoms of methane form a tetrahedron; the carbon atom lies in the center. *(image from LibreText core content.)*

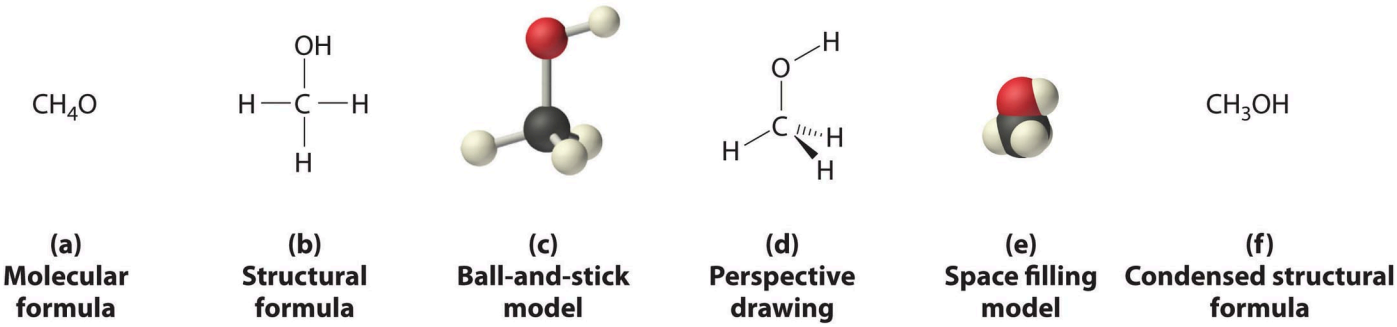


Figure 4.2.7: Different Ways of Representing the Structure of a methane (a) The molecular formula for methanol gives only the number of each kind of atom present. (b) The structural formula shows which atoms are connected. (c) The ball-and-stick model shows the atoms as spheres and the bonds as sticks. (d) A perspective drawing (also called a wedge-and-dash representation) attempts to show the three-dimensional structure of the molecule. (e) The space-filling model shows the atoms in the molecule but not the bonds. (f) The condensed structural formula is by far the easiest and most common way to represent a molecule. *(image from LibreText core content.)*

Although a structural formula, a ball-and-stick model, a perspective drawing, and a space-filling model provide a significant amount of information about the structure of a molecule, each requires time and effort. Consequently, chemists often use a condensed structural formula (part (f) in Figure 4.2.7), which omits the lines representing bonds between atoms and simply lists the atoms bonded to a given atom next to it. Multiple groups attached to the same atom are shown in parentheses, followed by a subscript that indicates the number of such groups. For example, the condensed structural formula for methanol is CH_3OH, which indicates that the molecule contains a CH_3 unit that looks like a fragment of methane (CH_4). Methanol can therefore be

viewed either as a methane molecule in which one hydrogen atom has been replaced by an –OH group or as a water molecule in which one hydrogen atom has been replaced by a $–CH_3$ fragment. Because of their ease of use and information content, we use condensed structural formulas for molecules throughout this text. Ball-and-stick models are used when needed to illustrate the three-dimensional structure of molecules, and space-filling models are used only when it is necessary to visualize the relative sizes of atoms or molecules to understand an important point.

Writing Condensed Molecular Formula

Test Yourself

Homework: Section 4.2

Exercise 4.2.1

Write the molecular formula of each compound. These answers explain the strategy.

a. The phosphorus-sulfur compound that is responsible for the ignition of so-called strike anywhere matches has 4 phosphorus atoms and 3 sulfur atoms per molecule.
b. Ethyl alcohol, the alcohol of alcoholic beverages, has 1 oxygen atom, 2 carbon atoms, and 6 hydrogen atoms per molecule.
c. Freon-11, once widely used in automobile air conditioners and implicated in damage to the ozone layer, has 1 carbon atom, 3 chlorine atoms, and 1 fluorine atom per molecule.

Strategy

- Identify the symbol for each element in the molecule. Identify the compound is an organic or an inorganic compound.
- For inorganic compounds, list the elements beginning with the one farthest left in the periodic table. List elements in the same group starting with the lower element and working up. For organic compounds, list carbon first, the hydrogen, then any other element in alphabetical order.
- From the information given, add a subscript for each kind of atom to write the molecular formula.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 4.2.2

Write the molecular formula for each compound.

a. Nitrous oxide, also called “laughing gas,” has 2 nitrogen atoms and 1 oxygen atom per molecule. Nitrous oxide is used as a mild anesthetic for minor surgery and as the propellant in cans of whipped cream.
b. Sucrose, also known as cane sugar, has 12 carbon atoms, 11 oxygen atoms, and 22 hydrogen atoms.
c. Sulfur hexafluoride, a gas used to pressurize “unpressurized” tennis balls and as a coolant in nuclear reactors, has 6 fluorine atoms and 1 sulfur atom per molecule.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 4.2.3

Write the molecular formula for each compound. The condensed structural formula is given.

a. Sulfur monochloride (also called disulfur dichloride) is a vile-smelling, corrosive yellow liquid used in the production of synthetic rubber. Its condensed structural formula is ClSSCl.
b. Ethylene glycol is the major ingredient in antifreeze. Its condensed structural formula is $HOCH_2CH_2OH$.
c. Trimethylamine is one of the substances responsible for the smell of spoiled fish. Its condensed structural formula is $(CH_3)_3N$.

Strategy:

- Identify every element in the condensed structural formula and then determine whether the compound is organic or inorganic.
- As appropriate, use either organic or inorganic convention to list the elements. Then add appropriate subscripts to indicate the number of atoms of each element present in the molecular formula.

Answer a available at the end of this section.

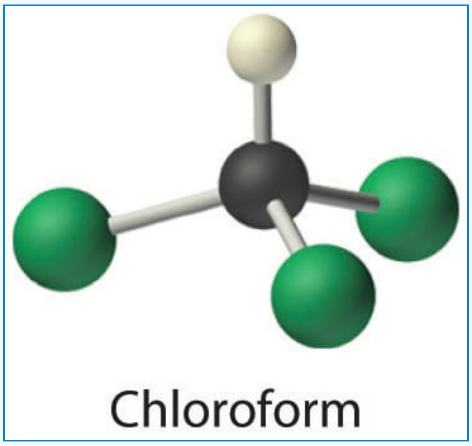
Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 4.2.4

Write the molecular formula for each molecule.

a. Chloroform, which was one of the first anesthetics and was used in many cough syrups until recently, contains one carbon atom, one hydrogen atom, and three chlorine atoms.



b. Hydrazine is used as a propellant in the attitude jets of the space shuttle. Its condensed structural formula is H_2NNH_2.

c. Putrescine is a pungent-smelling compound first isolated from extracts of rotting meat. Its condensed structural formula is $H_2NCH_2CH_2CH_2CH_2NH_2$. This is often written as $H_2N(CH_2)_4NH_2$ to indicate that there are four CH_2 fragments linked together.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

r

Answers

The molecule has 4 phosphorus atoms and 3 sulfur atoms. Because the compound does not contain mostly carbon and hydrogen, it is inorganic. Phosphorus is in group 15, and sulfur is in group 16. Because phosphorus is to the left of sulfur, it is written first. Writing the number of each kind of atom as a right-hand subscript gives P_4S_3 as the molecular formula.

Ethyl alcohol contains predominantly carbon and hydrogen, so it is an organic compound. The formula for an organic compound is written with the number of carbon atoms first, the number of hydrogen atoms next, and the other atoms in alphabetical order: CHO. **C** Adding subscripts gives the molecular formula]C_2H_6O.

Freon-11 contains carbon, chlorine, and fluorine. It can be viewed as either an inorganic compound or an organic compound (in which fluorine has replaced hydrogen). The formula for Freon-11 can therefore be written using either of the two conventions. According to the convention for inorganic compounds, carbon is written first because it is farther left in the periodic table. Fluorine and chlorine are in the same group, so they are listed beginning with the lower element and working up: CClF. Adding subscripts gives the molecular formula CCl_3F. We obtain the same formula for Freon-11 using the convention for organic compounds. The number of carbon atoms is written first, followed by the number of hydrogen atoms (zero) and then the other elements in alphabetical order, also giving CCl_3F

N_2O

$C_{12}H_{22}O_{11}$

SF_6

Each molecule of sulfur monochloride has two sulfur atoms and two chlorine atoms. Because it does not contain mostly carbon and hydrogen, it is an inorganic compound. Sulfur lies to the left of chlorine in the periodic table, so it is written first in the formula. Adding subscripts gives the molecular formula S_2Cl_2.

Counting the atoms in ethylene glycol, we get six hydrogen atoms, two carbon atoms, and two oxygen atoms per molecule. The compound consists mostly of carbon and hydrogen atoms, so it is organic. As with all organic compounds, C and H are written first in the molecular formula. Adding appropriate subscripts gives the molecular formula $C_2H_6O_2$.

The condensed structural formula shows that trimethylamine contains three CH_3 units, so we have one nitrogen atom, three carbon atoms, and nine hydrogen atoms per molecule. Because trimethylamine contains mostly carbon and hydrogen, it is an organic compound. According to the convention for organic compounds, C and H are written first, giving the molecular formula C_3H_9N.

$CHCl_3$

N_2H_4

$C_4H_{12}N_2$

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4.3: Ionic Compounds and Formulas

Learning objectives

- Recognize that metals lose electrons to form cations and that nonmetals gain electrons to form anions
- Predict the charge of monatomic main group elements based on their group number.
- Write formulas for ionic compounds using monatomic and polyatomic ions by applying the principle of charge neutrality.

Ionic Bonds and Ionic Compounds

If the number of protons in an atom or molecule is not equal to the number of electrons it has a net charge and is called an ion. Note, if it is an atom, it is called a monatomic ion, and if it is a molecule, it is called a polyatomic ion. There are obviously two types of ions, positive and negative, which are called cation and anion respectively. The charge of the ion is the number of protons minus the number of electrons.

Cation: Net positive charge, the number of protons is greater than the number of electrons
Anion: Net negative charge, the number of protons is less than the number of electrons.

Formula for Ionic Compounds.

When writing the formula of an ionic compound we use the lowest whole number ratio of cations ([+] ions) to anions ([-] ions). If you look at a salt crystal you will see that each positive ion is surrounded on all 6 sides by negative ions, and that a crystal lattice forms. So it makes no sense to have a "molecular formula" for an ionic compound, because the number of atoms will grow as the crystal grows.

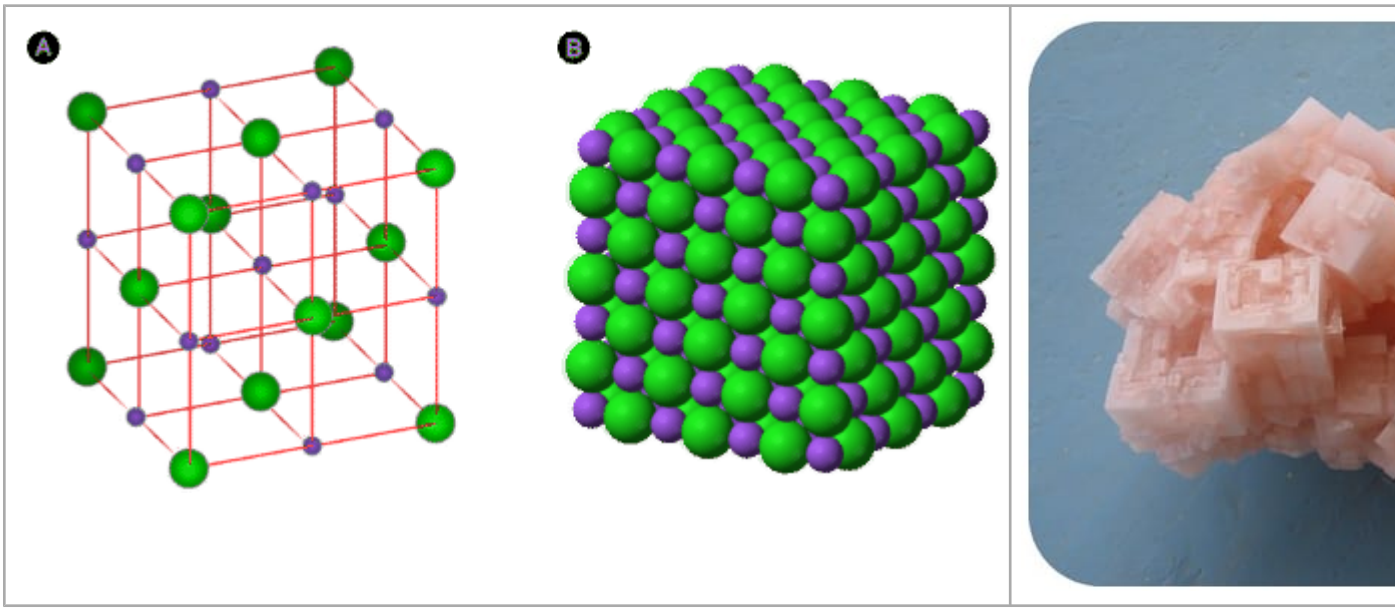


Figure 4.3.1 On the left is a crystal lattice of NaCl table salt, which is also shown on the right, where it forms Halite crystals. Because the number of ions depends on the size of the crystal, we use the lowest ratio of positive to negative ions that results in a neutral structure.

The concept of ionic bonds will be further developed in the next section.

Introduction

In the preceding section we discussed covalent bonds of molecules and introduced the concept of ionic bonds. In this section we will develop a deeper understanding of ionic compounds. We will see that to be stable an ionic compound must be neutral, and that there is not a discrete formula as there is with covalent compounds, in fact we use the lowest ratio of positive to negative charged ions that results in a neutral formula. In the previous section we did introduce the concept of cations (positive ions) and anions (negative ions). In the next section we will go over nomenclature, the naming of compounds.

Crystal Lattice and Principle of Charge Neutrality

Why must ionic compounds be neutral?

An ionic bond is not between two ions, but the result of the interaction of multiple cations ([+] ions) and anions ([-] ions) as the coordinate with each other and form an ionic crysta. Salts are typical ionic compounds. That is, a salt crystal has +/+ and -/- repulsions (increasing the energy), and +/- attractions (decreasing the energy, and making the compound more stable), and if the attractive forces are greater than the repulsive, an ionic compound is formed. In the following diagram we see how a crystal forms which maximizes the attractive interactions of the opposite charges while minimizing the repulsive interactions of the like charges. If the compound was not neutral, but had more cations, or more anions, the like/like repulsive forces would be larger and the material would become unstable and break apart. Therefore, ionic compounds must be neutral.

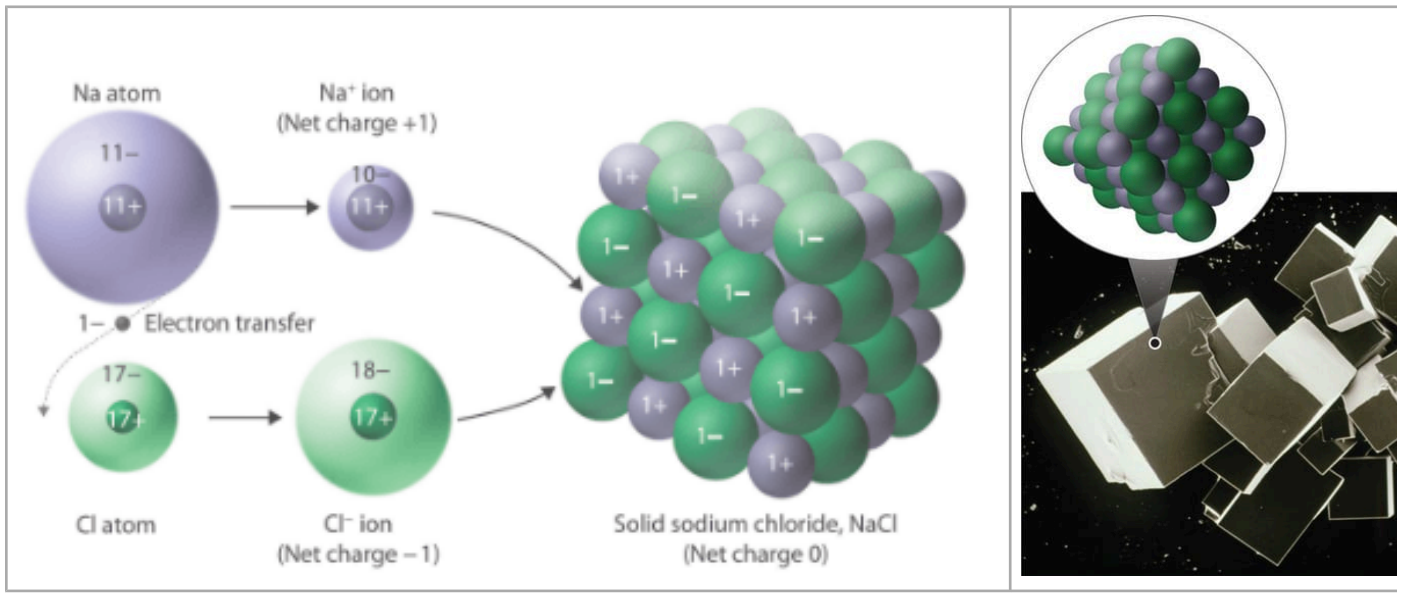


Figure 4.3.2: In the above image the sodium (a metal) loses and electron (becoming smaller) while the Chlorine atom (nonmetal) gains an electron (becoming bigger). A crystal lattice results where the cations and anions pack in three dimensional space that can result in a macroscopic crystal that is visible to the naked eye (CC BY-SA-NC; anonymous).

The salt crystals seen to the naked eye are the result of the way cations and anions pack to form a stable crystal structure that minimizes like/like repulsions and maximizes opposite charge attractions as defined by Coulomb's Law. The following video from YouTube describes the formation of sodium chloride.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 4.3.1: 2:10 min YouTube video describing formation of sodium chloride, table salt. Uploaded by Salvatore Lombardo, https://youtu.be/c-nMmwfTV-w.

Properties of Ionic Compounds

Because of the arrangement of cations and anions in a the crystal lattice, these ions are held tightly into their position. Although the ions can vibrate and oscillate about their position, it take much more energy than room temperature for one ion to break free of the lattice structure. This results in very high melting points and boiling points for ionic compounds. Therefore, ionic compounds are crystal solids at room temperature. They are not malleable or soft. When struck sharply, they do shatter cleanly, with layers of ions breaking off at once.

Monatomic Ions

Metals tend to form cations and nonmetals tend to form anions. The following periodic table shows some of the common ions formed by common elements. Note, the noble gasses (turquoise) do not tend to form ions.

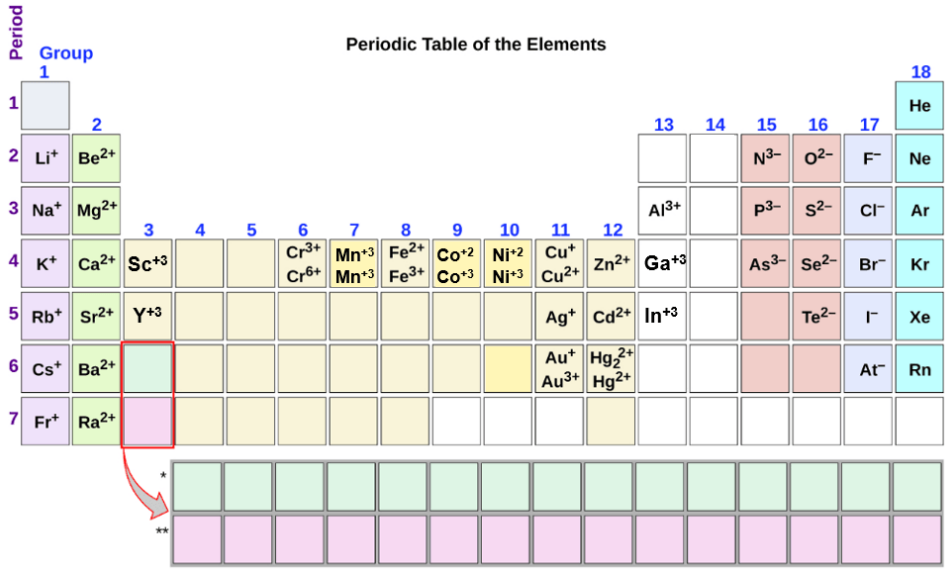


Figure 4.3.3: Note some metals form more than one type of monatomic cation (this table is not complete), and Hg_2^{2+} is not really a monatomic ion, but a diatomic ion.

Figure 4.3.3 should not be considered to be a comprehensive Figure and this topic will be brought up again in the next section. There are several things you can note from this table though. First, metals tend to from cations and nonmetals tend to form anions, while the noble gasses do not tend to form ions. Second, the nonmetals have only one charge state, which represents the number of electrons that need to be added for them to have the same number as a noble gas. Third, there are two types of metals, those that have only one charge state (like the nonmetals), and those that have multiple charge states. These different types of ion forming metals will be discussed more in the next section on naming compounds.

Monatomic Cations:

When atoms (typically metals) lose one or more electrons the become positively charged cations. For example, when a neutral sodium atom loses one electron, it becomes the sodium ion, Na^+ as shown in the following shorthand notation:

$$\mathrm{Na} \longrightarrow \mathrm{Na}^{+} + \mathrm{e}^{-} \tag{4.3.1}$$

Another example shows the formation of the magnesium ion from the neutral magnesium atom losing 2 electrons to pick up a +2 charge.

$$\mathrm{Mg} \longrightarrow \mathrm{Mg}^{2+} + 2\,\mathrm{e}^{-} \tag{4.3.2}$$

Can we predict the stable charge of a monoatomic cation?

Yes, it is easy for non-transition metals that have only one charged state (alkali metals, alkaline earth - see Figure 2.4.2). The trick is to remove enough electrons so that the cation has the same number of outer shell electrons as a noble gas, that is, it is isoelectronic to the noble gas. This is easy if we use the atom's group number's in terms of the "A" designation (remember there are two conventions for numbering the groups), as the group number effectively tells you how many electrons are needed to be removed. So Sodium (Group 1A) has 11 electrons and so if it looses one it has 10, like Neon (a noble gas). Magnesium has 12, so it looses 2 to become like Neon, and Aluminum has 13 and so looses 3 to become like Neon.

Note: Why does hydrogen not form a +1 cation?

Although hydrogen has one electron in its outer shell (like all the alkali metals), it only has one electron and so completely removing it would leave behind a bare nucleus, a subatomic particle the proton, which would not be stable. Consequently hydrogen does not form a cation, but can result in an acid, which is a polar covalent bond. In the above video sodium with 11 electrons gave an electron to chlorine, forming the Na^+ cation and the Cl^- anion, which then formed the crystal table salt NaCl. If chlorine was to try and remove the electron from hydrogen, it would not be able to as the resulting nucleus would pull it back, this results in the molecular acid HCl, where the chlorine can not completely remove the electron from the hydrogen. This is called a "polar covalent bond" and will be discussed later this semester.

Monatomic Anions:

*W*hen atoms (typically non-metals) gain one or more electrons they become negatively charged anions. For example, when a neutral fluorine atom gains one electrons, it becomes the fluoride ion as shown by the following shorthand notation:

$$\mathrm{F} + \mathrm{e}^{-} \longrightarrow \mathrm{F}^{-} \tag{4.3.3}$$

Another example is the formation of the sulfide ion when a neutral sulfur atom gains two electrons picking up a -2 charge.

$$\mathrm{S} + 2\,\mathrm{e}^{-} \longrightarrow \mathrm{S}^{2-} \tag{4.3.4}$$

Can we predict the stable charge of a monoatomic anion?
Yes, just as with the monatomic cations, stable monatomic anions are isoelectronic to the nearest noble gas, but in this case you add electrons instead of subtract. Once again, if we use the "A" designations for numbering the groups, we simply add the number of electrons required to be the same as a noble gas. So chlorine gains 1, oxygen gains 2 and nitrogen gains 3. This number can be determined by taking the group number 5, 6 or 7 from groups 5A, 6A and 7A and subtracting from 8 (for group 8A).

Exercise $4.3.2$

Based on the position of the periodic table, predict the charge of the following elements if they are to produce ions. (do not look at periodic table above to come up with your answer)

a. Sr
b. Ga
c. P

Answer available at the end of this section.

Ionic Compound Formula and the Principle of Charge Neutrality

The Principle of Charge Neutrality simply states that for matter to be stable it must be neutral. If an ionic compound had a formula of Na_2F, and you had 10,000 of those units, then there would be 20,000 positive things and 10,000 negative things, meaning an excess of 10,000 positive things repelling each other, and so it would not be stable, and so would not form (and realize that real substances on macrosocopic scales that we can measure have in the order of 10^{23} entities.

You might ask, if something is neutral, why are the ion-ion attractions greater than the repulsion, that is, why isn't the energy of interaction neutral? The reason that the attractions are greater than the repulsions is because opposite ions tend to be closer together (they are attracted) than the ions of same sign (they are repulsed), and according to Coulomb's law, the closer they are the greater the force, and so a neutral compound has a net attractive force and forms a stable compound.

The formula of an ionic compound represents the lowest whole number ratio of cations to anions, it is as simple as that. Most cations have charges of [+1] through [+6] while most ions have charges of [-1] through [-3].

Trick:
Set # of Anions = Charge of Cation and
set # of Cations = Charge of Anion
Beware that this must be the Lowest Whole # Ratio of cation to anion.

Strategy:

1. **Identify charge of each ion**
2. **multiply charge of each ion by absolute value of the charge of the counter ion.**
3. **Make sure resulting ratio is the lowest whole number ratio, if not, divide by common denominator so all values are integers.**

Watch video 4.3.2 to see how to perform this calculation

YouTube Video

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View this YouTube Video online

Video 4.3.2: (6'15") Video on using Principle of Charge Neutrality to determine formula of Ionic Compound

Exercise 4.3.3

What is the formula for the ionic compound that forms when the calcium ion reacts with the oxide ion?

Answer available at the end of this section.

s

Exercise 4.3.4

What is the formula for the ionic that forms when the lead(IV) (Pb^{4+}) ion reacts with the sulfate (SO_4^{2-}) ion?

Answer available at the end of this section.

Determining Ion Charges from Formula

If given a formula like NaCl or FeO or Fe_2O_3 you need to be able to calculate the charge of the ions. In the case of NaCl it is easy as Na is always [+1] and Cl is always [-1]. But iron can have multiple charges and so needs to be Figured out.

Fortunately, the anion is always of known charge and so you can use that to algebraically reverse engineer the principle of charge neutrality. that is for a general formula C_xA_y where C is the Cation with charge [+m] and there are x of them, and A is the Anion with charge of [-n] and there are y of them, we know from the principle of charge neutrality that

$$\begin{gathered} x(+m)+y(-n)=0 \\ \therefore \\ x=\frac{0-y(-n)}{x} \end{gathered} \tag{4.3.5}$$

So for FeO we know the charge of the oxygen is -2, so 1(charge of Fe) +1(charge of O) = 0, so Fe=+2 as)=-1

for Fe_2O_3 we have

2(charge of Fe) + 3(charge of Oxygen) = 0

so

$$\begin{gathered} 2(+m)+3(-2)=0 \\ \therefore \\ +m=\frac{0-3(-2)}{2}=3 \end{gathered} \tag{4.3.6}$$

that is: 3x2 -2x3=0, which makes sense.

Exercise 4.3.5

What is the the charge for the cation in each of the following comounds?

a. V_2O_5

b. Pb_3N_4

Answer available at the end of this section.

Test Yourself

Homework: Section 4.3

Graded Assignment: Section 4.3

Query 4.3.1

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Contributors

- Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Elena Lisitsyna (H5P interactive modules)

Answers

A) $\mathrm{Sr^{2+}}$. Sr is a group 2A main group metal. So, +2 charge.

B) $\mathrm{Ga^{3+}}$. Ga is a group 3A main group metal. So, +3 charge.

C) $\mathrm{P^{3-}}$. P is a group 5A main group nonmetal. So, 5-8 = -3 charge.

CaO. Ca^{2+} and O^{2-}. You need one calcium ion for every one oxide ion for the compound to be charge neutral.

Add $\mathrm{Pb(SO_4)_2}$. $\mathrm{Pb^{4+}}$ and $\mathrm{{SO_4}^{2-}}$. For every one Lead(IV) ion, you need two sulfate ions to be charge neutral. 1(+4) + 2(-2)=0. You may use the criss-cross rule, but you must make sure that you reduce to the simplest whole number ratio.

a. V^5 because oxygen is minus 2.

b. Pb^4 because nitrogen is minus 3.

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4.4: Polyatomic Ions

Polyatomic Ions

Polyatomic Ions are charged molecules, where the total number of protons in all bonded atoms is $\neq$ to the total number of electrons in all bonded electrons. When writing structural formulas for ions we place the formula in brackets and the charge as a superscript outside of the bracket. The number of protons is the total number of protons in all atoms, while the number of protons minus the number of electrons is the charge of the atom. That is the charge of the ion is the charge of all protons [+ value] plus the charge of all electrons [- value].

- Ammonium is a polyatomic cation, NH_4^+.

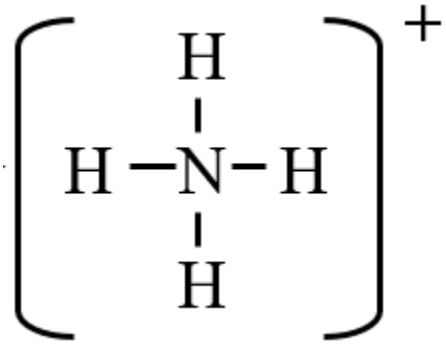


Figure 4.4.1: Structural formula for ammonia, NH_4^+. Note how the formula is placed in a bracket with the charge outside of the bracket. You must always indicate the charge of the ion when writing it (Copyright; author via source)

Exercise 4.4.1

How many protons and electrons are in ammonia

Answer available at the end of this section.

- Sulfate is a polyatomic anion, SO_4^{-2}.

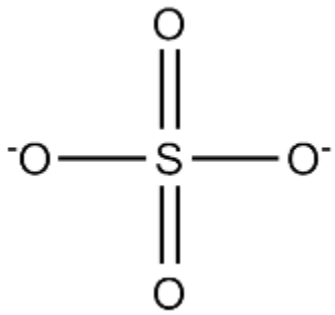


Figure 4.4.2: Structural formula for sulfate, SO_4^{-2}. Note how the formula is placed in a bracket with the charge outside of the bracket. You must always indicate the charge of the ion when writing it.

Exercise 4.4.1

How many protons and electrons are in sulfate?

Answer available at the end of this section.

Formula of Polyatomic Ions

The rules for formula of polyatoic ions are the same as for monatomic ions. The formula of the neutral compound between ammonium and sulfate is $(NH_4)_2SO_4$.

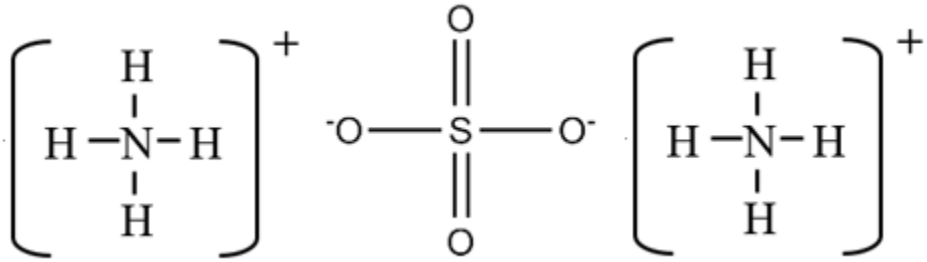


Figure 4.4.3: There are two ammonium ions for every sulfate ion.

Answers

Exercise 4.4.1

There are 7 protons from the nitrogen and 4 from the four hydrogen (1 each) and so a total of 11 protons are in this polyatomic ion, which as drawn, has 4 covalent bonds. Since the charge is [+1],

$$\#protons - \#electrons = charge$$
$$\#electrons = \#protons - charge$$
$$\#electrons = 11 - 1 = 10$$

Exercise 4.4.1

There are 16 protons from the sulfur and 32 from the four oxygen (8 each) and so a total of 48 protons are in this polyatomic ion, which as drawn, has 6 covalent bonds (note two are double bonds). Since the charge is [-2],

$$\#protons - \#electrons = charge$$
$$\#electrons = \#protons - charge$$
$$\#electrons = 48 - (-2) = 50$$

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4.5: Nomenclature

Learning Objectives

1. Recognize ionic compound, molecular compounds, and acids.
2. Give the names and formulas for ionic compounds
3. Give the names and formulas for molecular compound s
4. Give the name and formula for acidic compound

Introduction

Nomenclature, the naming of chemical compounds is of critical importance to the practice of chemistry, as a chemical can not only have many different names, but different chemicals can have the same name! So how can you do science if you do not know what you are talking about??? In this section we will look at nomenclature of simple chemical compounds. The rules we use depends on the type of compound we are attempting to name.

Overview of Nomenclature

The rules we use in nomenclature depend on the types of bonds the compound has and these are outlined in Figure 4.5.1 and the video lecture in video 4.5.1. That is, before we do anything we have to ask what kind of compound are we trying to name?

- The first question we ask is if the compound is ionic or covalent?
  - **If it is covalent**, which is typically between 2 or more nonmetals, we need to ask, is it a simple molecule, or is it an acid.
    - If it is **a simple molecule we use Greek prefixes** to identify the number of atoms of each type of element in the molecule.
    - If it is **an acid, we base it's name on the ionic compound it would form if hydrogen could be a cation**. Note, hydrogen can not lose its only electron as then it would be a subatomic particle and the charge density would be too high, so it forms a covalent bond.
  - **If it is ionic**, we **use the principle of charge neutrality to name the compound**.

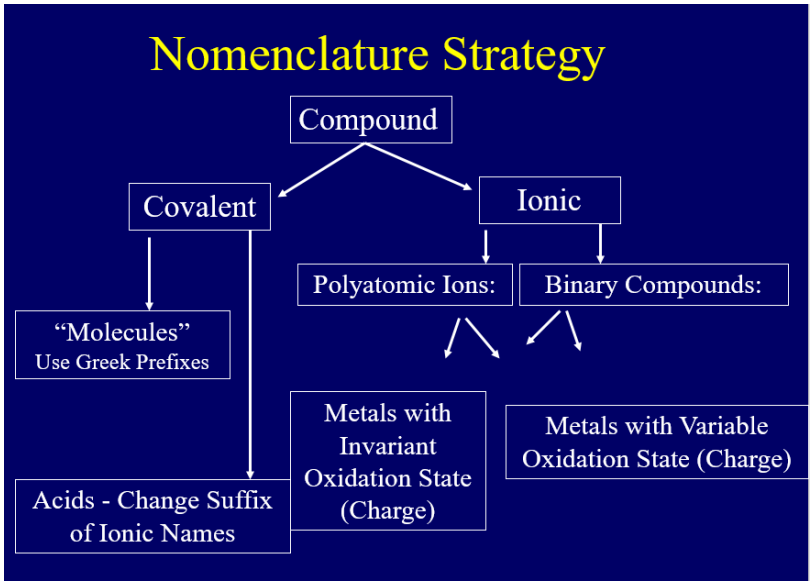


Figure 4.5.1: Outline of strategy for naming molecules (watch the Youtube immediately below for a lesson on how to use this outline)

YouTube Video

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View this YouTube Video online

Video 4.5.1: YouTube (2'17") describing the logic of the flow diagram for (https://youtu.be/4Ikom9GjAhE)

The following lesson will start with naming ionic compounds, then naming covalent compounds and then acids. After that, we will look at some special types of compounds, hydrated compounds (which are ionic compounds that contain water in their structure) and acid salts, which are salt's with acidic anions.

Ionic Compounds

When naming an ions the cation uses the metal's elemental name followed by the word ion or cation and the nonmetal anion uses the elemental name with an ide ending. So the metal sodium (Na) forms the sodium ion or sodium cation (Na^+), while the nonmetal chloride (which exists as Cl_2) is called chloride (Cl^-), and you do not need to indicate that it is an ion (there is nothing wrong with the terms chloride anion or chloride ion, but they are providing redundant information).

Exercise 4.5.1

Give the symbol and name for the ion with 34 protons and 36 electrons. Is it a cation or an anion?

Answer available at the end of this section.

Ionic compounds have a [+] cation and a [-] anion. We use the **Principle of Charge Neutrality,** that is, for an **ionic compound to be stable its chemical formula MUST BE NEUTRAL.** So **you do not need to state the number of cations and anions**, you **only need to state what they are** (and what the charge is for cations with variable charges). You then Figure the formula based on the lowest whole number ratio of cations to anions that produces a neutral formula.

We will start with binary ionic, which are between monatomic anions and cations. We note that there are two types of metals, those that have only 1 charge (Type 1), and those that can have more than one stable charge. Finally, polyatomic ions often form which are covalently bonded atoms where the total number of protons is not equal to the total number of electrons.

Binary Ionic

Binary ionic compounds are between a metal and nonmetal. This does not mean there are two atoms, but two types of atoms, so Al_2S_3 is a binary ionic compound. The rules are simple, name the cation first and the anion second, giving the anion the -ide suffix.

- **Cation (metal) first**
- **Anion (nonmetal) second with -ide suffix**

The problem is some metals have more than one stable charge state, and so for those metals (type 2 metals), we need to identify the charge in the name.

Two Types of Monatomic Cations

There are two types of monatomic cations, those of metals that form only one charge state, and those that can form multiple charged states, but there is only one type of monatomic anion as none of them form multiple charged states (review section

2.6.4.1) For simplicity, we are going to label the metal cations as Type 1 (invariant charge) and type 2 (variable charge)

Type 1: Invariant Charge (Oxidation State)

Remember metals lose electrons to form cations and nonmetals gain electrons to form anions. Figure 4.5.2 lists the ions (cations and anions) that have invariant charges. All the anions are of this type, gaining the number of electrons required to fill the valence shell and become isoelectronic with the nearest noble gas.

There is a pattern to the metals if you use the A/B block notation in that 1A & 1B tend to form [+1] , IIA & IIB tend to form [+2], while IIIA & IIIB tend to form [+3]. It is easy to see how the alkaline metals (1A) lose one electron as that gives them the same as a noble gas, but you need to get to Chapter 7 and understand the orbitals before you can understand why silver (1B) behaves the same. The alkaline earths (IIA) each lose 2 and gain the same electron configuration of a noble gas, while Zinc and Cadmium (IIB) also lose two electrons, but once again, we need to get to Chapter 7 to understand why. As for the [+3] forming ions, these are IIIA and IIIB. Now the B block are easier to see, as they simply loose three electrons, one from the IIIB, IIA and IA block of a period, and then they have the same number of electrons as a noble gas. As for the IIIA, we need to understand Chapter 7, but imagine that all of the orbitals of the transition metals are filled (B block) and the ones lost are from IA, IIA and IIIA.

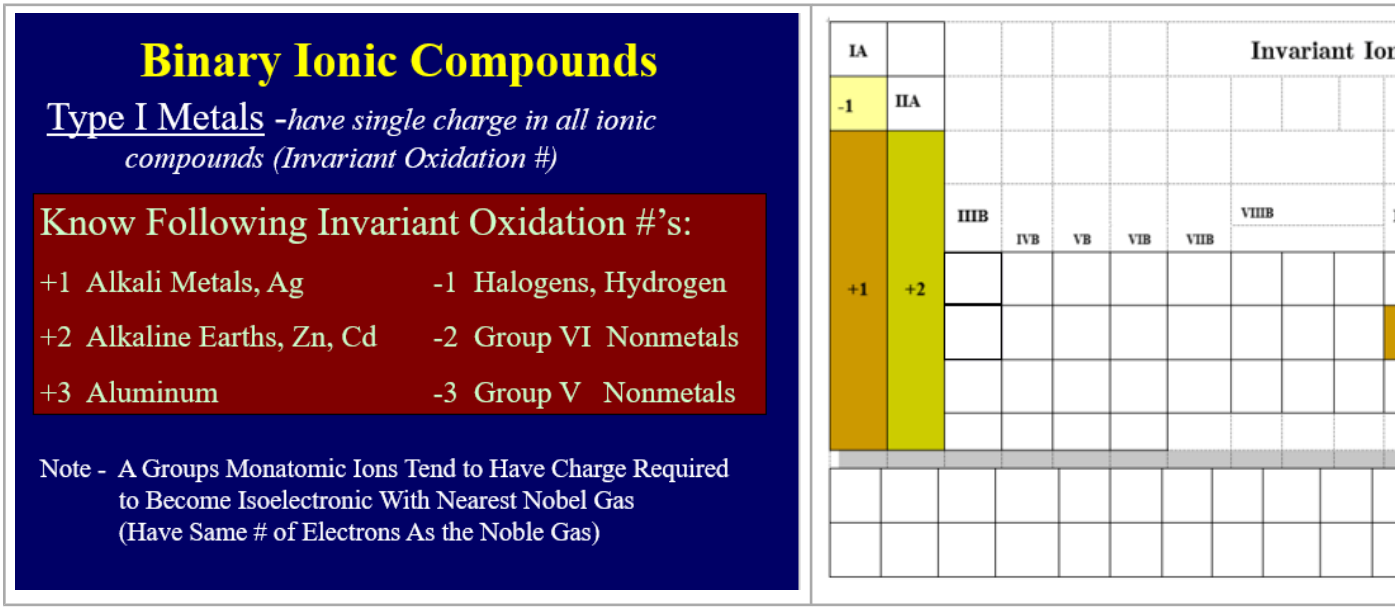


Figure 4.5.2 The elements in the above table form only one stable charged state. Can you see the pattern for the cations in relation to the A/B block notation. Note, we often refer to the charge of a monatomic ion as its oxidation state, and the list of Type 1 Metals (invariant oxidation state) is more extensive in general chemistry (section 2.7.3.1)

YouTube Video

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Video 4.5.2: 3:10 min YouTube on nomenclature for binary ionic compounds (with type I metal cations) This video uses the list from general chemistry, prep chem students need to recognize the invariant oxidation state metals in figure 4.5.2..

Using the Principle of Charge Neutrality (section 4.1.2) and knowing the charge of the ions allows you to determine the formula. For type 1 metals is is very simple, you just name the metal, then the nonmetal and give the nonmetal an -ide ending. So table salt (NaCl) is sodium chloride.

Exercise 4.5.2

What is the formula for Barium Chloride?

Answer available at the end of this section.

Exercise 4.5.3

Name $AlBr_3$

Answer available at the end of this section.

Type II: Cations with Variable Charge (Oxidation State)

Most transition metals form multiple stable cations with different charges, and so you have to identify the charge state, which is done by writing the charge in Roman numerals and placing it in parenthesis after the name of the metal. So Fe^{+2} is iron(II) and Fe^{+3} is iron(III). Some of the metals form very common ions which have latin names that are in common use, and general chemistry students need to be familiar with those in the following table but prep chem do not.

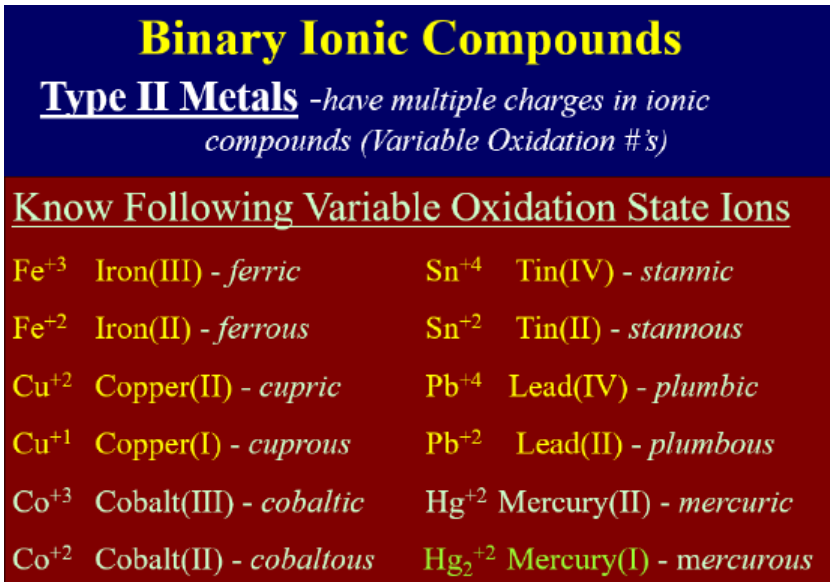


Figure 4.5.3: Some common type 2 metal and their stable ions.

Note the latin convention is to use:

- ***-ous*** ending for the **lower** oxidation state (charge)
- ***-ic*** ending for the **higher** oxidation state (charge)

Note **mercury(1) is not a monoatomic cation**, but is really a diatomic cation of two mercury atoms bound to each other, both having lost one electron. Polyatomic ions will be covered in the next section

Exercise 4.5.4

What is the formula of mercury(I) chloride?

Answer available at the end of this section.

Example 4.5.1

Give the names and formula of the two chloride salts that from with the common cations of iron.

Solution

NOTE: Review section 2.6.5 and Video 2.6.2 if you have problems calculating the formula in this example or the next exercise.

Example 2

Ions:	Fe^{2+} and Cl^-	Fe^{3+} and Cl^-
Compound:	$FeCl_2$	$FeCl_3$
Nomenclature	Iron (II) Chloride or Ferrous Chloride	Iron (III) Chloride or Ferric Chloride

Exercise 4.5.5

Give the names and formula of the two nitride salts that from with the common cations of lead.

Answer available at the end of this section.

Polyatomic Ions

Polyatomic ions are covalent units (molecules) where the total number of protons $\neq$ to the total number of electrons and they were introducedin section 4.4. They end in [-ide], [-ate] and [-ite]. Many of these are oxyanions with oxygen being bonded to a nonmetal and others are carboxylate ions . It should be noted that ionic compounds like ammonium acetate are composed entirely of nonmetals, but they form a crystal structure of cations and anion.

Table 4.5.1: List of Polyatomic ions freshmen chemistry students are required to know

Formula	Name		Formula	Name
NH_4^+	Ammonium		SO_3^{2-}	Sulfite
Hg_2^{2+}	Mercury (I) ion		SO_4^{2-}	Sulfate
ClO^-	Hypochlorite		$S_2O_3^{2-}$	Thiosulfate
ClO_2^-	Chlorite		CrO_4^{2-}	Chromate
ClO_3^-	Chlorate		$Cr_2O_7^{2-}$	Dichromate
ClO_4^-	Perchlorate		CN^-	Cyanide
NO_2^-	Nitrite		SCN^-	Thiocyanate
NO_3^-	Nitrate		OH^-	Hydroxide
CO_3^{2-}	Carbonate		O_2^{2-}	Peroxide
HCO_3^-	Bicarbonate		MnO_4^-	Permanganate
PO_4^{3-}	Phosphate		$C_2O_4^{2-}$	Oxalate
PO_3^{3-}	Phosphite		$C_8H_4O_4^{2-}$	Phthalate
			$C_2H_3O_2^-$	Acetate

Naming Oxyanions: Oxyanions are polyatomic ions where oxygen is attached to a nonmetal, nonmetals of the same periodic group from homologous oxyanions. Lets look at the 4 oxyanions of bromine

- **per**brom**ate** (BrO_4^-)
- brom**ate** (BrO_3^-)
- brom**ite** (BrO_2^-)
- **hypo**brom**ite** (BrO^-)

These all end with -ate or -ite, and the real reasoning deals with the oxidation state of the nonmetal, sort of its hypothetical charge within the covalently bonded polyatomic ion. We will cover that in a later Chapter, but right now we note that the "-ate" ions have more oxygens and the "-ites" have less, with "per___ate" having the most, and "hypo__ite" having the least.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 4.5.3: YouTube video going over how to remember the names, formulas and charges of oxyanions (https://www.youtube.com/embed/AOtrIEjZd4c).

Overview of video:

- Oxyacids have same charge as central atom's monatomic ion (except for the first period)
  - nitride, phosphide,arsenide = [-3] so oxyanions of phosphorous & Arsenic = [-3], but of nitrogen = [-1]*
  - oxide, sulfide, selenide, telluride = [-2] so oxyanions of sulfur, selenium & tellurium = [-2], oxygen does not form same kind of anions
  - fluoride, chloride, bromide, iodide = [-1], so oxyanions of chlorine, bromine, & iodine = [-1], fluorine's also = [-1] but I think only hypofluorite really exits, although I have seen the others on exams. For example, a search of PubChem for perfluorate gives no hits.
- ate implies more oxygens
  - hyper__ate implies even more oxygen
- ite implies less oxygens
  - hypo__ite implies even fewer oxygen.

*Note, nitrite and nitrate are the only oxyanions with a charge different than the monatomic anoin

Naming Carboxylate ions: Carboxylate ions are very common in organic chemistry. The following YouTube uses the structure of the carboxylate functional group as a tool to assist in remembering the formulas.

YouTube Video

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View this YouTube Video online

Video 4.5.4: Nomenclature of oxalate, phthalate and acetate, all of which have the carboxylate functional group.

Exercise 4.5.6

Name the following compounds:

A) $NiCO_3$

B) $Sr(NO_2)_2$

C) NH_4NO_3

Answer available at the end of this section.

Covalent Compounds

Molecules are compounds held together by covalent bonds. In figure 4.5.1 we see that there are two types of covalent compounds, what we are calling "classical molecules", and acids. Hydrogen like all the alkali metals only has one electron in its outer shell, but unlike the other elements of the first period (group 1A) you can not remove the electron and make an ion because all that is left would be a proton, which is a sub-atomic particle. This creates a class of molecules called acids, where hydrogen effectively takes the place of the cation, and the resulting bond is a polar covalent bond where the electron is unequally shared. Classical molecules occur when nonmetals form covalent bonds with each other and this can get very complicated. Organic chemistry is based on the study of carbon based molecules and uses IUPAC nomenclature which is beyond the level of this course (the IUPAC Nomenclature book for organic compounds is 1568 pages long!). We will first cover simple molecules and then acids.

Simple Molecules

In this class we will look at the nomenclature of very simple small molecules that form when two or more nonmetals share covalent electrons and form a classical molecule. In a molecule the formula represents the chemical entity, where as in an ionic compound, it is the lowest whole number ratio of cation to anion. So molecules like P_2O_5 and P_4O_{10} have the same ratio of phosphorous to oxygen, but they are different molecules, and so we need a nomenclature system that tells how many atoms of each element are in the compound.

The rules are simple, before each atom we place a Greek prefix indicating the number of those atoms in the formula. If there is only on atom, we often omit the prefix "mono". In this class we will name molecules with no more than 10 atoms of each type and you need to memorize the 10 prefixes in table 4.5.2.

Table 4.5.2: Common greek prefixes used in naming simple moleculs

# of Atoms	1	2	3	4	5	6	7	8	9	10
Prefixes	Mono-	Di-	Tri-	Tetra-	Penta-	Hexa-	Hepta-	Octa-	Nona-	Deca-

There are some rules we need to follow,which are:

1. If there are two atoms, place the more "metallic" first (furthest to the left on the periodic table)
2. is there is a central atom, place it first, and the atoms attached after.
3. If there are more than two atoms, place in order of bonds

Example 4.5.2

Give the names for: CO, BCl_3

Give the formula for: carbon dioxide, dinitrogen pentoxide

Solution

CO = carbon ***mon***oxide BCl_3 = boron***tri***chloride

carbon dioxide= CO_2 dinitrogen pentoxide = N_2O_5

Exercise 4.5.7

Name each molecule.

a. SF_4
b. P_2S_5
c. N_2O_4
d. P_4S_{10}

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Exercise 4.5.8

Give the formula for each molecule.

a. disulfur difluoride
b. iodine pentabromide
c. nitrogen trifluoride
d. diphosphorous pentaoxide

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Acids

In a binary ionic compound an alkali metal typically loses an electron to a nonmetal forming an [+1] cation while the nonmetal forms an anion. The resulting cation has core (inner shell) electrons that give it volume and surround the nucleus with electrons resulting in a stable ion, even if the net charge is positive. Hydrogen is also in group IA but has no inner shell electrons, and so it can't form a cation, as that would be a subatomic particle and the charge density would be so great that it would attract the electron back. So what happens is an unequal sharing of the electron and hydrogen forms an acidic bond with anions that is of the same formula as the other group 1A alkali metals would form. So acid nomenclature is based on that of an ionic compound.

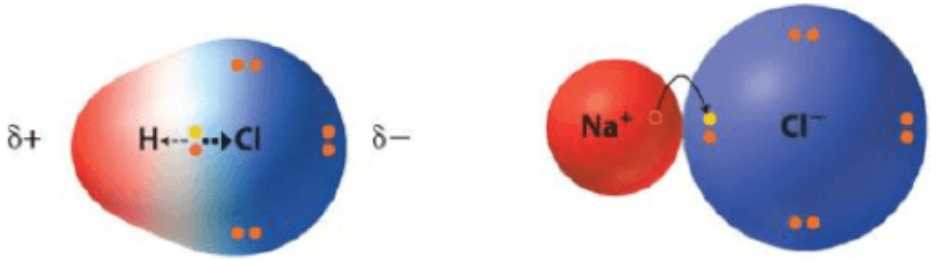


Figure 4.5.4: Hydrogen forms acids of the same "formula" as other alkali metals. On the left is the acid with chloride, HCl and on the right the salt of chloride NaCl.

When an acid dissolves in water the proton is transferred to a water forming hydronium ion H_3O^+ and thus dissociates forming the same anion as the analogous salt.

$$\begin{aligned} HCl(aq) &\rightarrow H^+ + Cl^- \\ NaCl(aq) &\rightarrow Na^+ + Cl^- \end{aligned} \tag{4.5.1}$$

where H^+ really exists as H_3O^+ in the water (the proton is transferred to water).

So we could say sodium chloride is a salt with a common ion to HCl. Because of this, we base the name of the acid on it's salt. Since there are three endings to salts (-ide, -ite, -ate), there are three correlating names for acids.

- **-ide** becomes **hydro** [anion stem]**ic acid**
- **-ite** becomes [anion stem]**ous acid**
- **-ate** becomes [anion stem]**ic acid**

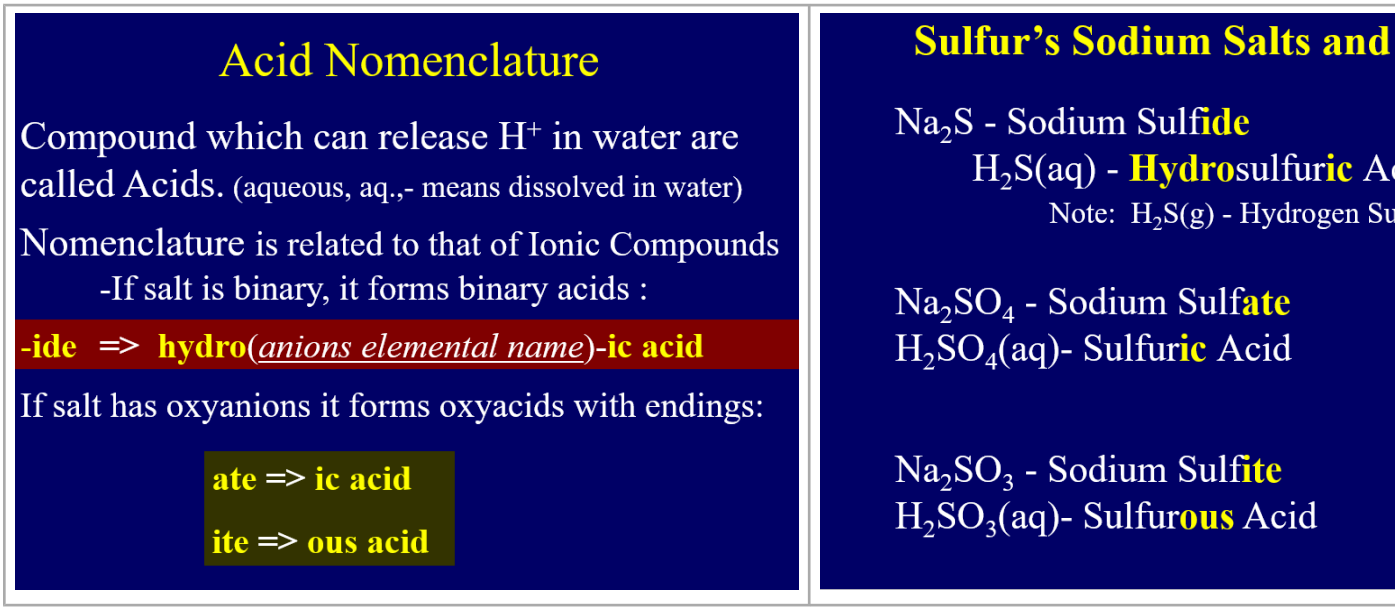


Figure 4.5.5: The rules for acid nomenclature as based on the nomenclature of ionic compounds

Compare the questions and answers of exercise 4.5.9 and 4.5.10 to so this relationship for the salts and acids of chlorine

Exercise 4.5.9

Name the five sodium salts of chlorine.

a. NaCl
b. NaClO
c. $NaClO_2$
d. NaClO3
e. NaClO4

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Exercise 4.5.10

Name the five acids of chlorine.

a. HCl
b. HClO
c. $HClO_2$
d. HClO3
e. HClO4

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Test Yourself

Homework: Section 4.5

Graded Assignment: Section 4.5

Query 4.5.1

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View this Interactive Content online

The following set has 24 questions and randomly loads 10

Query 4.5.2

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Query 4.5.3

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The following set has 31 questions and randomly generates 10

Query 4.5.4

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- November Palmer, Ronia Kattoum & Emily Choate (UALR)
- Elena Lisitsyna (H5P interactive modules)

Answers

Se^{2-}, selenide. In it an anion since there are more electrons than protons.

Barium is an alkaline earth and always forms a cation of charge of [+2], while chlorine is a halogen and always form the chloride ion of [-1]. For barium chloride to be neutral you would need two chlorides for every barium, and so the formula is $BaCl_2$.

Aluminum is always [+3] and bromide is always [-1] so it is **aluminum bromide**. The name does not have to tell how many cations and anions there are, as we know there are three [-1] bromides for every one [+3] aluminum ion, as otherwise it would not be neutral (review section 2.6.5.1)

Hg_2Cl_2, for every one Hg_2^{-2} cation there must be two Cl^- anions. So Hg_2Cl_2 it the lowest whole number ratio of cation to anion.

Ions:	Pb^{2+} and N^{-3}	Pb^{4+} and N^{-3}
Compound:	Pb_3N_2	Fe_3N_4
Nomenclature	Lead (II) nitride or Plumbous nitride	Lead (IV) nitride or Plumbic Nitride

A) **Nickel(II) carbonate**. Nickel is a Type II cation, so you need a Roman Numeral. To Figure out that charge "x", we need take a look at what we know. There is one carbonate ion at a charge of -2. There is one Nickel ion at unknown charge "x". 1(x) + 1(-2) = 0. To be charge neutral, x=+2.

B) **Strontium nitrite.** Strontium is a type I cation and there is no need for a Roman Numeral. NO_2^- is the nitrite ion.

C) **Ammonium nitrate.** both are polyatomic ions, so you simply state the name of cation followed by the name of the anion.

sulfur tetrafluoride

diphosphorus pentasulfide

dinitrogen tetraoxide

tetraphosphorus decasulfide

S_2F_2

IBr_5

NF_3

P_2O_5

Sodium chloride

Sodium hypochlorite

Sodium chlorite

Sodium chlorate

Sodium perchlorate

Hydrochloric acid

Hypochlorous acid

Chlorous acid

Chloric acid

Perchloric acid

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5: CHEMICAL REACTIONS



- Topic hierarchy

5.1: Introduction to Chemical Reactions

5.2: Types of Chemical Reactions

5.3: Balancing Chemical Equations

5.4: Aqueous Reactions

5.5: Neutralization Reactions

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5.1: Introduction to Chemical Reactions

Learning Objectives

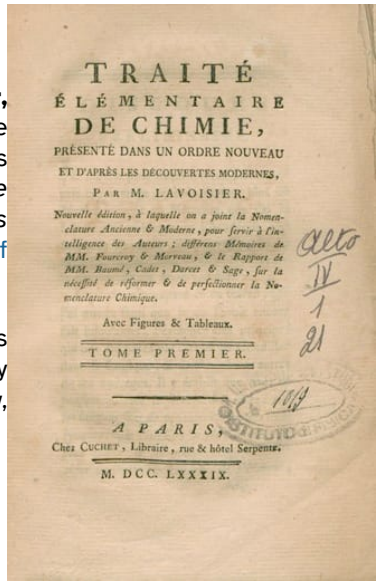
- Define terms used to represent chemical reactions including reactants, products, states, coefficients, and stoichiometry, reaction conditions.
- Apply the law of conservation of matter to balanced chemical reactions
- Differentiate between the liquid phase and the aqueous phase

What are Chemical Equations?

Chemical equations are how chemists describe chemical reactions, the process by which one form of matter (reactants) turn into another form of matter (products). Below is the unbalanced chemical reaction for the combustion of methane (CH_4) with oxygen (the reactants) to produce carbon dioxide and water (the products).

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$
(unbalanced equation)

Law of Conservation of Mass: Balancing Reactions


TRAITÉ
ÉLÉMENTAIRE
DE CHIMIE,
PRÉSENTÉ DANS UN ORDRE NOUVEAU
ET D'APRÈS LES DÉCOUVERTES MODERNES,
PAR M. LAVOISIER.
Nouvelle édition, à laquelle on a joint la Nomenclature Ancienne & Moderne, pour servir à l'intelligence des Auteurs ; différens Mémoires de MM. Fourcroy & Morveau, & le Rapport de MM. Baumé, Cadet, Darcet & Sage, sur la nécessité de réformer & de perfectionner la Nomenclature Chimique.
Avec Figures & Tableaux.
TOME PREMIER.
A PARIS,
Chez CUCHET, Libraire, rue & hôtel Serpente.
M. DCC. LXXXIX.

Antoine Lavoisier (1743-1794) is attributed to introducing the **Law of Conversation of Matter,** which states that during a chemical process mass is neither created or destroyed, and thus the total mass of the products equals the total mass of the reactants. This was first published in his 1789 "Traité Élémentaire de Chimie" and in France this was known as Lavoisier's Law, "*Rien ne se perd, rien ne se crée, tout se transforme. (Nothing is lost, nothing is created, everything is transformed.*"). The Gutenberg Project offers a free English translation of the book (Elements of Chemistry)..

As matter is made of compounds and elements, and the compounds may change, this means the total number of atoms of each element must be conserved. In the above equation only atoms of carbon are conserved, and so the equation is not balanced. In the equation below, that atoms of each element are conserved, and so it is balanced.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$
(balanced equation)

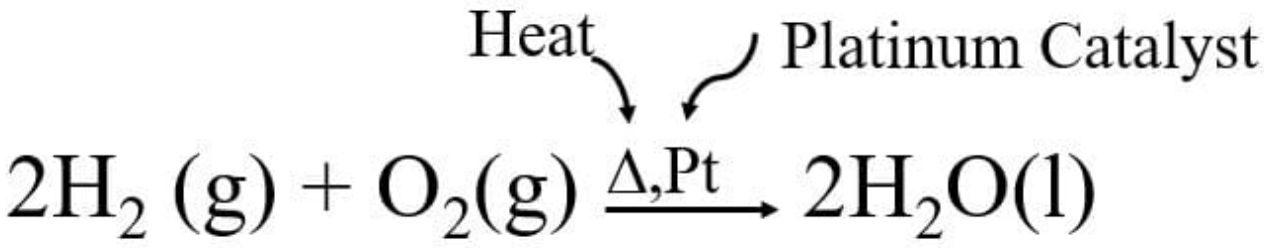
Stoichiometric Coefficients:

Once the reactants and products are written, you are not able to change the subscripts of the compound to balance the chemical reaction without changing the identity of the reactants and products. However, you may use stoichiometric coefficients in front of each chemical entity (molecule or element) to balance the equations. The word "stoichiometry" come from the Greek *stoicheion* (element) and *metrein* (to measure).

What other information can be included in a chemical equation?

- Phases
  - (s) - Solid
  - (l) - liquid
  - (g),(v) - gas, vapor
  - (aq) - aqueous
- Reaction Conditions (often indicated above the arrow)

- (Δ) - Heat
- Catalyst



Aqueous vs. Liquid Phase:

Liquid is a phase, the condensed fluid phase (in contrast to a gas, which is the dispersed fluid phase). Aqueous is a solution where something is dissolved in water, the solvent. NaCl(l) is molten sodium chloride, that is obtained by heating it above the melting point. NaCl(aq) is aqueous sodium chloride, which is made by dissolving the salt into water. At one atm pressure and normal temperatures, you can have NaCl(aq) (salt water), but you can't have molten NaCl (l), because the melting point is 1,474^{o}F.

Catalysts:

Catalysts can be understood by looking at the mechanism of a reaction, that is, how it actually proceeds, and can be considered to be chemicals that function as both a reactant and a product. That is, the catalyst reacts with something and form a new chemical, an intermediate, and that intermediate also reacts, and reproduces the catalyst. Catalysts can speed up reactions, or make reactions happen which without the catalyst are so slow that they don't really happen.

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Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Ronia Kattoum

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5.2: Types of Chemical Reactions

learning objectives

- List and define the 5 types of reactions based on the rearrangement of atoms
- Match given reactions to their respective category of reaction type
- Relate combustion reactions of hydrocarbons to climate change
- Relate decomposition reactions to production of hydrogen fuel cells.

Some Basic Types of Chemical Reactions

We are going to identify 5 basic "types" of chemical reactions because they have characteristic equations that allow us to use different techniques when we balance them. That is, in the next section when we balance a reaction, the first thing we will ask is what type of reaction it is, and then we will use the technique that is appropriate to that type reaction. There are other ways of classifying reactions that may actually tell us more about the chemistry, like are electrons transferred between atoms when a reaction occurs, which we will discuss later in this Chapter, and are called oxidation-reduction (redox) reactions, or acid base reactions (which at this level, occur when protons are transferred between between reactants).

1. **Formation Reactions (synthesis reactions)**
2. **Decomposition Reactions**
3. **Single Displacement Reactions**
4. **Double Displacement Reaction**
5. **Combustion Reactions**

Formation or Synthesis Reactions

$$H_2(g) + Cl_2(g) \text{ ® } 2HCl(g)$$

As the name implies, two or more chemicals come together and form another chemical. Here, HCl is being formed from hydrogen and chlorine.

Decomposition Reaction

$$CaCO_3(s) \text{ --> } CaCO(s) + CO_2(g)$$

As the name implies, something decomposes into something else. Here Calcium Carbonate decomposes to calcium oxide and carbon dioxide

Single Displacement Reactions

$$Fe(s) + CuSO_4(aq) \text{ -> } FeSO_4(aq) + Cu(s)$$
$$F_2(g) + BaCl_2(aq) \text{ -> } BaF_2(g) + Cl_2(g)$$

In a single displacement reaction something that is "pure" displaces something in a compound. As a rule of thumb, we say metals form cations and nonmetals from anions. So a metal can exist in two forms, as a pure metal, or as a cation (in an ionic compound). Likewise, a nonmetal can exist as something pure, or as an anion (in an ionic compound). In the first of the above two equations, the metal iron displaces the cation Copper(II) and we call that a metal single displacement reaction, while in the second, the fluorine displaces the anion chloride, and so we call that a nonmetal single displacement reaction. Note, sometimes these are called single replacement reactions.

If you note, these reactions involve a change in the charge of the metal or nonmetal, and so they involve the transfer of electrons. We will pick these up again later in this chapter, but a single displacement reactions involves the transference of electrons from

the element that was pure to the element that was a cation, causing the cation to become pure, and the pure element to become a cation. If we just looked at the metals in the above equation we note that

$$Fe(s) + Cu^{+2}(aq) \rightarrow Fe^{+2}(aq) + Cu(s)$$

In this class we will focus on single displacement reactions involving metals, but in general chemistry you will need to understand these for both metals and nonmetals.

Metal Single Displacement

$$Fe(s) + AuSO_4(aq) \rightarrow FeSO_4(aq) + Au(s)$$

Some metals prefer to be ions, and other prefer to be pure. This can be understood by comparing iron ore with gold ore. Iron (pure) spontaneously rusts and iron ore that is mined must consists of compounds like hematite (Fe_2O_3) and magnetite (Fe_3O_4), while gold, a noble metal, does not tend to form ionic compounds, and so pure gold can be found in gold ore. So the above reaction is logical, in the sense that iron as a pure metal is not stable, while gold is.



Figure 5.2.1: Hematite ore on left (image from flickr) and gold on right (image from wikicommons).

Since the above reaction proceeds, we can correctly predict that:

$$FeSO_4(aq) + Au(s) \rightarrow \text{No Reaction}$$

Metal Displacing Nonmetal

A metal can also displace hydrogen from an acid or even from water (look at water as H-OH, where the anion is hydroxide). For example

$$2Na(s) + 2H_2O(l) \rightarrow H_2(g) + 2NaOH$$

Here the pure sodium metal becomes a cation, while a hydrogen is pulled off of each water to from hydrogen gas (a diatomic), and so hydrogen is behaving similar to a cation on the reactant side, but it is not a cation, it is an acidic hydrogen atom in a covalent bond. These reactions can give off heat which can then combust the hydrogen, as seen in Video 5.2.2.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.2.1. 1'09" YouTube uploaded by halosaqi showing alkali metals displacing hydrogen from water.

Double Displacement Reactions

$NiCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + NiCO_3(s)$

In a double displacement reaction, the ions of two ionic compounds swap counter ions. In this equation a precipitate formed

Nickle(II)chloride and sodium carbonate

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.2.2: double displace reaction producing a preipitate

We will look at these in great detail in sections 5.3 and 5.4

Combustion Reaction

In a combustion reaction a compound reacts with oxygen and releases energy. Carbon dioxide and water are also released if the compound contains carbon.

$CH_4 + 2O_2 \dashrightarrow CO_2 + 2H_2O$ + energy

Do combustion reactions always produce carbon dioxide and water?

No there are other types of combustion reactions, like the combustion of hydrogen and magnesium.

$2H_2 + O_2 \dashrightarrow 2H_2O$ + energy
$2Mg + O_2 \dashrightarrow MgO$ + energy

The hydrogen economy is considered as an alternative to fossil fuels and an area of intense current chemistry research. The concept is simple, one can use electricity generated by methods like solar, hydroelectric, wind, geothermal to split water into hydrogen and oxygen. Then you can transport the hydrogen to a new location and recombine them to produce energy where you need it, and produce no carbon dioxide. There are obvious challenges, like the storage of the hydrogen. As we shall learn later this semester, energy is the capacity to do work and transfer heat. So any energy lost as heat can not be used for work. You have probably heard of fuel cells, and they operate like a battery that consumes fuel, and can be more efficient than a combustion engine because they do not produce as much heat. In the following video the person uses a reversible fuel cell kit to demonstrate the concept of a hyrdogen economy. Simply speaking, you split water into hydrogen and oxygen when you are near a power

source (solar, geothermal, wind, hydroelectric) then you recombine them where you want to use the energy, and in the process, you produce no carbon dioxide.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.2.3: 4'18" YouTube Uploaded by Andrew Copeland showing basic concepts behind hydrogen economy as an alternative to the fossil fuel economy (https://youtu.be/qZhhnBfXP7I).

Practice Problems

Exercise 5.2.1

Identify the types of reactions for each equation

a. $AlCl_3(aq) + Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + NaCl(aq)$
b. $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$
c. $Li_2S + Br_2 \rightarrow 2LiBr + S$
d. $H_2CO_3 \rightarrow H_2O + CO_2$
e. $Ba(C_2H_3O_2)_2(aq) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + 2HC_2H_3O_2(aq)$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Exercise 5.2.2

Identify the types of reactions for each equation

a. $H_2S(g) \rightarrow H_2(g) + S(s)$
b. $Pb + 2CuCl \rightarrow PbCl_2 + 2Cu$
c. $CH_2CO(g) + 2\ O_2(g) \rightarrow 2\ CO_2(g) + H_2O(g)$
d. $HCl + NaOH \rightarrow NaCl + H_2O$

Answer a available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Predicting Products

For single displacement, double displacement and combustion reactions you need to be able to predict the products from the reactants.

Single Displacement

One of the reactant species is a pure elemental substance and the other is involved in a compound. If it is a metal displacement reaction the the cation in the compound becomes pure and the metal becomes a cation.

$$Metal + \underbrace{Cation_aAnion_b}_{\text{(neutral formula of ionic compound)}} \rightarrow Metal \quad \underbrace{Cation_xAnion_y}_{\text{(neutral formula of ionic compound)}} \tag{5.2.1}$$

$$Al + FeSO_4 \rightarrow ? \tag{5.2.2}$$

Aluminum only has one ion, Al^{+3} and so the product salt is the neutral salt between Al^{+3} and SO_4^{-2}, $Al_2(SO_4)_3$.

$$\underset{\text{(unbalanced equation)}}{Al + FeSO_4 \rightarrow \ Fe + Al_2(SO_4)_3} \tag{5.2.3}$$

Exercise 5.2.3

$$Al + Fe_2(SO_4)_3 \rightarrow ? \tag{5.2.4}$$

Answer available at the end of this section.

Double Displacement

For double displacement you swap counterions.

$$\underbrace{Cation_aAnion_b}_{\text{(neutral formula of ionic compound)}} + \underbrace{Cation_cAnion_d}_{\text{(neutral formula of ionic compound)}} \rightarrow \underbrace{Cation_wAnion_x}_{\text{(neutral formula of ionic compound)}} + \underbrace{Cation_yAnion_z}_{\text{(neutral formula of ionic compound)}} \tag{5.2.5}$$

Strategy

1.Identify Reactant Ions and their Charge
2. Swap Ions
3. Use Principle of Charge Neutrality to determine formula of products

Example 5.2.1

$$Al_2(SO_4)_3 + NH_4ClO_4 \rightarrow ? \tag{5.2.6}$$

Solution

1. Identify reactant ions and their charge

Al^{+3}	SO_4^{-2}
NH_4^+	ClO_4^-

Table 5.2.1: Identifying ions and their charges

2. **Swap Ions**

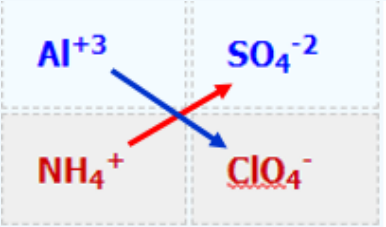


Table 5.2.2: Swapping Ions

3. Using Principle of Charge Neutrality to Identify Products

$$Al_2(SO_4)_3 + NH_4ClO_4 \rightarrow Al(ClO_4)_3 + (NH_4)_2SO_4 \quad (5.2.7)$$
(unbalanced equation)

Combustion

For carbon based compounds the products of complete combustion are carbon dioxide and water.

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O \quad (5.2.8)$$

(students need to know the products for combustion of carbon based compounds)

Test Yourself

Homework: Section 5.2

Graded Assignment: Section 5.2

Query 5.2.1

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Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- November Palmer & Emily Chaote (UALR)
- Vincent Belford (H5P)
- Ronia Kattoum (Learning Goals)

Answers

Double Displacement

combustion

Single Displacement (non metal)

Decomposition

Double Displacement (involving acid)

Decomposition

Single Displacement (metal displacement)

Combustion

Double Displacement (neutralization)

Exercise 5.2.3

$$Al + Fe_2(SO_4)_3 \rightarrow \ Fe + Al_2(SO_4)_3 \quad (5.2.9)$$
(unbalanced equation)

Note, iron is type II metal and forms more than one stable ion, while Aluminum is type one and is always [+3], and so the products are the same as the above illustration.

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5.3: Balancing Chemical Equations

learning objective

- Balance chemical reactions based on reaction type

The first step is to identify reactants and products. Sometimes these will be given in the problem statement, and for certain types of problems, you will need to be able to predict the products.

Basic techniques for balancing equations

1. **Identify reactants and products**
   - Be sure to write correct formulas
   - Never change the subscripts in the formulas when balancing an equation
   - Read the question carefully. Problem statements may say "when molecule A is formed from the reaction of molecule B and C", and students will treat molecule A as a reactant, when it is really a product.
2. **Write Unbalanced Skeletal Equation**
   - Write Reactants to left of arrow.
   - Write Products to right of arrow.
   - Leave a space for stoichiometric coefficient.
     - placing a line in front of each chemical is a good way to do this

'_' allowed only in math mode

3. **Identify Type of Reaction**
   1. Single Displacement
      - Balance ions not atoms (Video 3.3a)
   2. Double Displacement
      - Balance ions not atoms (Video 3.3b)
      - Strategy is pick an ion, do its counterion, then do the counterion's counterion, and repeat, until all have been balanced (watch video).
   3. Combustion
      - Balance Carbon first, hydrogen second and oxygen last
   4. All other
      - Balance by inspection
      - If an element exists by itself, like Fe, O_2, S_8, balance it last.
4. **Check Work**

Single Displacement Reactions

$$Al(s) + FeSO_4(aq) \rightarrow ?$$

The trick is to balance the ions. In the following video, we do not balance Al, Fe , S and O, but aluminum and sulfate. You know from the previous section that the aluminum displaces the iron in the sulfate salt, and so you also need to be able to predict the products (aluminum and iron sulfate form iron plus aluminum sulfate)

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.3.1: 44" video uploaded by Bob Belford showing how quickly you can balance a single displacement reaction by balancing the ions, not the atoms. (https://youtu.be/KTHURNRwd5Y)

Double Displacement Reaction

$$Al_2(SO_4)_3 + NH_4ClO_4 \rightarrow ? \tag{5.3.1}$$

In the double displacement reaction the ions swap partners. As long as the same ions exist on both sides of the equation (none convert to another compound or ion) you can simply balance the ions, not the atoms. So in this video we balance aluminum, sulfate, ammonium and perchlorate, and not Al, S, O, N, H and Cl.

YouTube Video

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View this YouTube Video online

Video 5.3.2: 1'12" youtube uploaded by Bob Belford on balancing double displacement reactions. Note the "path" used to go from ion pair to ion pair.

Combustion Reactions

Fire, the combustion reaction dates to the dawn of humanity. We typically think of fire as the combustion of organic compounds as they react with oxygen to produce carbon dioxide and water, but combustion is actually a broader class of reactions where compounds react with oxygen to produce oxides and release energy. In the following video you see magnesium metal combusts, not only in air, but in water.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.3.3: 1'26" video of the Royal Society of Chemistry showing a lecture demonstration with magnesium burning in air and water, (https://youtu.be/gAhsvOP5yaM).

In this class you need to know that carbon based compounds combust with oxygen and produce carbon dioxide and water. The

general unbalanced equation is

$$C_xH_yO_z \ + \ O_2 \rightarrow CO_2 \ + \ H_2O \tag{5.3.2}$$

In balancing carbon based combustion reactions there is a schema that makes things very easy.

1. Balance carbon
2. Balance hydrogen
3. Balance oxygen (you may need to use a fractional coefficient)
4. If you used a fractional coefficient, multiply by the denominator so all stoichiometric coefficients are integers.

The following video shows how to balance the combustion reaction for the hydrocarbon C_8H_{18}.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.3.4: 1'06" video balancing combustion reaction of octane. Note, Digital corrigendum: Octane is C_8H_{18} and I wrote C_8H_8 by mistake in the second step, but treated it as C_8H_{18} , and will redo when I have time, but the technique/answer is correct except for that typo (https://youtu.be/FfpRu2k23js).

The next video shows an even more complicate problem, the combustion of $C_9H_{17}O_3H$.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.3.5: 54" video showing the combustion of $C_9H_{17}O_3H$ (https://youtu.be/ulSV3b-MQRY).

Test Yourself

Homework: Section 5.3

Graded Assignment: Section 5.3

Exercise 5.3.1

Balance the following equations

a. $AlCl_3(aq) + Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + NaCl(aq)$
b. $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$
c. $Li_2S + Br_2 \rightarrow LiBr + S$
d. $C_8H_{18}O + O_2 \rightarrow CO_2 + H_2O$
e. $Ba(C_2H_3O_2)_2(aq) + HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + HC_2H_3O_2(aq)$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Contributors and Attributions

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- Ronia Kattoum (Learning Objectives)

Answers

$2AlCl_3(aq) + 3Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + 6NaCl(aq)$

$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

$Li_2S + Br_2 \rightarrow 2LiBr + S$

$C_8H_{18}O + 12O_2 \rightarrow 8CO_2 + 9H_2O$

$Ba(C_2H_3O_2)_2(aq) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + 2HC_2H_3O_2(aq)$

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5.4: Aqueous Reactions

Learning Objectives

- Outline the differences between strong electrolyte, weak electrolyte, and a nonelectrolyte
- Predict the solubility of ionic compounds in water using solubility rules
- Memorize seven strong acids that are strong electrolytes
- Distinguish ways of writing aqueous reaction equations
  - General (molecular) ionic
  - Total (complete) ionic
  - Net Ionic

The objective of this section is to predict what will happen for single and double replacement reactions that occur in aqueous solutions. Note, if in a double displacement reaction two solutions combine and form a solid, you have a precipitation reaction. Many texts treat this as a type of reaction, but we will treat it as a subset of the double displacement reactions.

Electrolytes

Pure water does not conduct electricity, and it has been observed that when a substance dissolves in water, it may produce mobile ions that allow the water to conduct electricity, and we call that compound an electrolyte, or it may not, in which case we call it a nonelectrolyte.

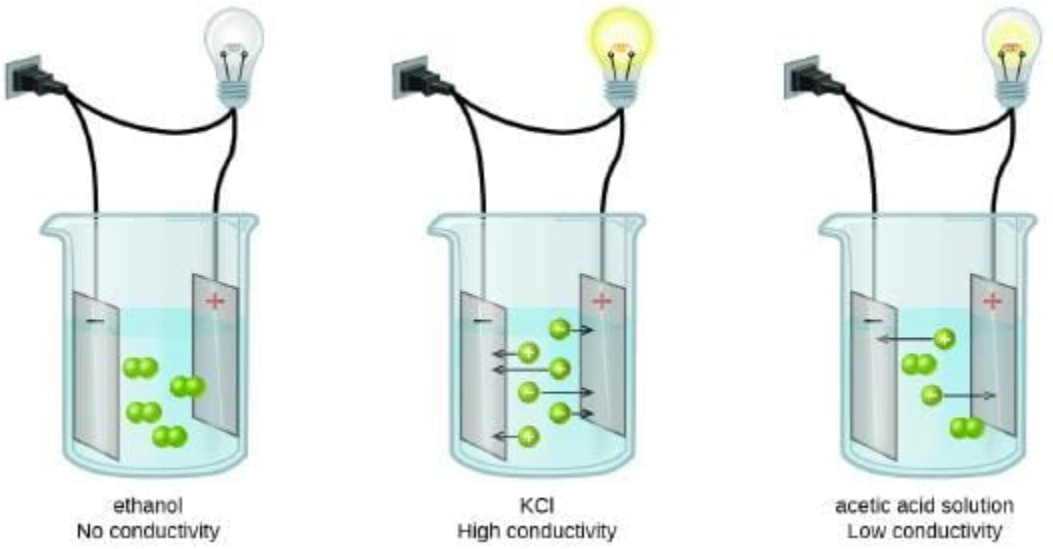


Figure 5.4.1: Ethanol on the left is a nonelectrolyte and does not conduct electricity. KCl is a strong electrolyte and the bulb is very bright. Acetic acid is a weak electrolyte, and although the image may not show it, if the concentrations are the same, the light is dimmer than for the KCl

Exercise 5.4.1

Both salt (NaCl) and table sugar, glucose ($C_6H_{12}O_6$) dissolve in water. Why is salt water a strong electrolyte, while sugar water is a non-electrolyte?

Answer available at the end of this section.

There are two basic ways an aqueous compound can be an electrolyte; being a soluble ionic compound or a strong acid.

Soluble Ionic Compounds

- If ionic compounds dissolve and form a solution, the ions separate and are free to move about and conduct the electricity. But not all ionic compounds dissolve, and so they can be weak, strong or even nonelectrolytes. Typically, the nondissolved ionic compound forms a solid that falls to the bottom as a precipitate.

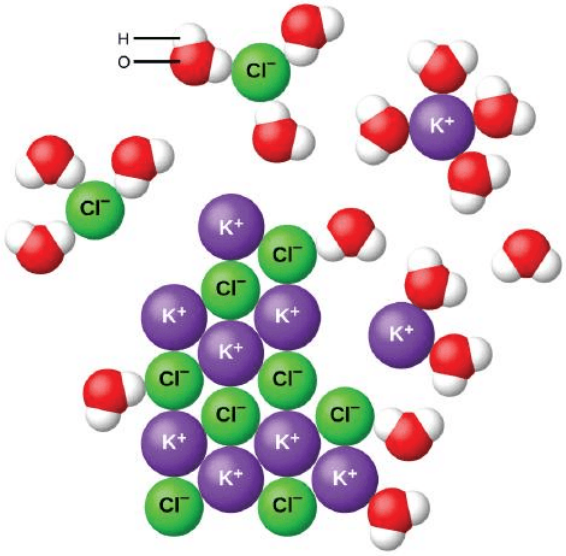


Figure 5.4.2: In the above image, the solid KCl is being surrounded by water molecules which cause the ions to leave the crystal and enter the solution. Once they enter the solution they are mobile and can conduct electricity, so KCl is an electrolyte (it is actually a strong electrolyte)

Types of Electrolytes

Compounds can be Strong, Weak, or Nonelectrolytes

1. **Strong Electrolytes** - Strong Conductors of Electricity due to formation of a large number of Mobile Ions
2. **Weak Electrolytes** -Weak Conductors of Electricity due to formation of a few Mobile Ions
3. **NonElectrolytes** –Non Conductors of Electricity as they do not form Ions in aqueous solutions

Strong Electrolytes

Ionic - Soluble Salts and Strong Bases

$NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

Covalent - Strong Acids (protonate water)

$HCl(aq) + H_2O \rightarrow H_3O^+(aq) + Cl^-(aq)$

$H_2SO_4(aq) + H_2O \rightarrow H_3O^+(aq) + HSO_4^-(aq)$

Weak Electrolytes

Ionic - Slightly Soluble Salts

$CoCl_2(s) \rightleftharpoons Co^{+2}(aq) + 2Cl^-(aq)$

Covalent - Weak Acids & Amine Bases (hydrolyze water)

$HF(aq) + H_2O \rightleftharpoons H_3O^+(aq) + F^-(aq)$

$NH_3(aq) + H_2O \rightleftharpoons NH_4^+(aq) + OH^-$

NonElectrolytes

Ionic - Insoluble Salts

$CoS(s) <=--> Co^{+2}(aq) + S^{-2}(aq)$

Covalent - Molecules which do not hydrolyze or protonate water

$C_{12}H_{22}O_{11}(s) + H_2O --> C_{12}H_{22}O_{11}(aq)$

The following animations gives an atomic scale visualization of strong, weak and nonelectrolytes

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.4.1: 2'37" YouTube animation uploaded by Georgianna Allen giving an explanation for how some compounds can be weak, strong or non-electrolytes (https://youtu.be/B5mSet29qxA).

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.4.2: 1'29" YouTube animation uploaded by BerkeleyChemDemos showing conductivity of various solutes. Note, the solid ionic compounds do not conduct because their ions are not mobile. You should pause the video before they add the light bulbs and try and predict if the bulb will come on.

Predicting Products for Double Displacement Reactions

Determine what the products will be for the following reactions. Initially we will ignore phases, but by the time we finish this exercise you will need to be able to identify phases as you write the equation. This process requires that you identify the ions in the reactants and for double displacement reactions swap partners and use the principle of charge neutrality to determine what the product are. For single replacement reaction you need to Figure which species (anion or cation) is gaining or losing charge and the other species is a spectator ion. Once that is done you need to balance the equations

Students frequently swap ions without thinking about the product formula and this leads to mistakes. It is suggested you follow these steps:

1. **Identify reactant ions and their charge**
2. **Swap Ions**
3. **Determine Product Formula based on principle of charge neutrality**
4. **Balance Equation**

Example:

$$Pb\left(ClO\right)_2\left(aq\right) + Na_2SO_4\left(aq\right) \rightarrow \tag{5.4.1}$$

1. **Identify reactant ions and their charge**

Pb^{+2}	ClO^-
Na^+	SO_4^{-2}

Table 5.4.1: Matrix showing the cations and ions for each of the two ionic compounds in the double displacement reaction

2. **Swap Ions**

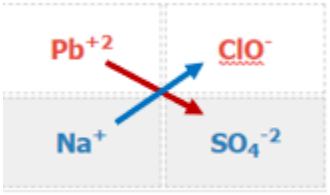


Table 5.4.2: Matrix of Table 5.4.1 showing the cations and ions swapping partners.

$$\mathbf{Pb^{+2} + SO_4^{-2}} \rightarrow ?$$

$$\mathbf{Na^{+} + ClO^{-}} \rightarrow ?$$

Note: If dealing with a strong acid, treat the "H" as a cation (break the acid HA into H^+ and A^-, where A is any anion, noting the number of hydrogens equals the charge of the anion, as the acid must be neutral). We will cover acids in the next section.

3. **Determine Product Formula based on principle of charge neutrality**

$$Pb\left(ClO\right)_2\left(aq\right) + Na_2SO_4\left(aq\right) \rightarrow PbSO_4\left(?\right) + NaClO\left(?\right) \tag{5.4.2}$$

If you need help determining formulas review section 2.6

4. **Balance Equation**

$$Pb\left(ClO\right)_2\left(aq\right) + Na_2SO_4\left(aq\right) \rightarrow PbSO_4\left(?\right) + 2NaClO\left(?\right) \tag{5.4.3}$$

Exercise 5.4.2

Predict products for the following reactions

a. $Mg_3\left(PO_4\right)_2 + NaClO_3 \rightarrow ?$
b. $AlPO_4 + Na_2SO_3 \rightarrow ?$

Answer a available at the end of this section.

The next step is is figuring out the phase of the product. There are 3 common phases for aqueous reactions, solid (precipitate), aqueous and gas. In the case of aqueous products resulting from mixing two solutions, we state there is no reaction. In general chemistry 2 we will use mathematical formula of solubility products, but in this class we will use the Solubility rules, which are a good "rule of thumb" to determine if a salt will dissolve or not. Now note, these are all relative values, and even a soluble salt will form a precipitate, if you add a lot of it. For example, you can dissolve around 140 grams of potassium iodide in 100 g of water at 32°F, which makes it a very soluble salt, but a precipitate will occur if you try and add 200 g of KI to 100 g of water.

Solubility Rules

We saw that some compounds are soluble, other are partially soluble, and some are completely insoluble. How do we know determine this? We will start off with the simplest types, which are ionic compounds, and we will base these on the nature of the ions. There are sort of two opposing processes going on. Are the attractions of the ions in the crystal stronger than their attraction towards the water, or weaker? If the ionic attractions within the crystal are stronger, they do not dissolve and they form a precipitate. If on the other hand, the ions are more attracted to the water, they leave the crystal and the compound is soluble. We will use the solubility rules to determine if a salt is soluble or not.

What are the Solubility Rules?

These are what I am calling a "rule of thumb," and allow us to roughly predict if a salt will dissolve or not. It must be understood that the concept is relative, for example, table salt is considered a soluble salt and if you add table salt to water it will dissolve a lot, up to 359g per liter, but at that point it becomes saturated, and any more will form a precipitate. On the other hand silver chloride is an insoluble salt, and you can only dissolve 0.0019g into 1 liter, but any more will fall to the bottom as a precipitate.

Note, some textbooks give slightly different rules, and this set is incomplete. If your text is different, please discuss this with your instructor. When you get to general chemistry 2 you will learn a different approach, where we can quantify the amount dissolved for an insoluble salt, like the 1.9 mg/liter for silver chloride.

Is there a strategy to using the solubility rules?
Yes, for the ones I have set up below.

- We first look at the [+] cation, and ask if it is any one of the cations listed in step IA. If yes, we say it is soluble, and the question is answered. After this step we focus on the [-] anions. If it is not soluble from step 1A, we go to 1B, and if the anion is from this list, it is soluble.
- We now go to the compounds that are usually soluble, step II, and you need to memorize the exceptions, which are insoluble.
- We now go to the compounds that are usually insoluble, step III, and you need to memorize the exceptions, which are soluble.

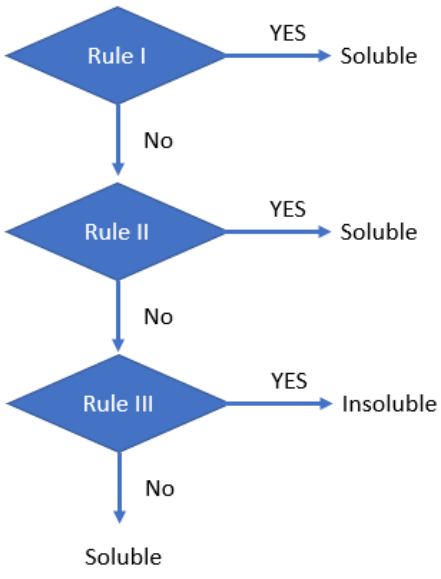
Solubility Rules

Table 5.4.3: Solubility rules with schema showing flow diagram. NOTE: the exceptions in steps II and III have opposite meanings.

I. Soluble
   a. Group 1A & Ammonium (the only cations in this list)
   b. NO_3^-, ClO_4^-, ClO_3^-, $CH_3CO_2^-$

II. Usually Soluble
   a. Cl^-, Br^-, I^-, (Except those with Ag^+, Hg_2^{+2}, & Pb^{+2})
   b. F- (Except Mg^{+2}, Ca^{+2}, Sr^{+2}, Ba^{+2} & Pb^{+2})
   c. SO_4^{-2} (Except those with Ca^{+2}, Sr^{+2}, Ba^{+2}, Ag^+ & Pb^{+2})

III. Insoluble (Except with cations from I.a)
   a. OH^- (Except those with Sr^{+2} & Ba^{+2})
   b. Everything else (this is not true, but will work in this class)



In the following video we will first predict the products for the double displacement reaction of aluminum sulfate and barium chloride, and then apply the solubility rules to determine the phases of the products.

$$Al_2(SO_4)_3 + BaCl_2 \rightarrow ?$$

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 5.4.3: YouTube predicting products and phases for a double displacement reaction (https://youtu.be/e1gzOZxT51w).

Exercise 5.4.3

Use the solubility rules to determine if the following compounds are soluble or insoluble. Indicate answer by writing formula followed by (aq) for soluble and (s) for insoluble. (You may also want to write the names of the species)

a. $PbSO_4$
b. $NaClO_4$

Answer a available at the end of this section.

Exercise 5.4.4

Identify the following as a strong electrolyte, a weak electrolyte, or a non-electroylte.

a. $Ba(NO_3)_2$
b. H_3PO_4
c. $C_6H_{12}O_6$
d. HNO_3
e. AgBr

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Exercise 5.4.5

Identify the following as a strong electrolyte, a weak electrolyte, or a non-electroylte.

a. $NH_4)_3PO_4(aq) + NaOH(aq) \rightarrow$
b. $(NH_4)_3PO_3(aq) + CaSO_4(s) \rightarrow$
c. $(NH_4)_2SO_4(aq) + CaCl_2(aq) \rightarrow$
d. $AgCH_3CO_2\ (aq) + KCl(aq) \rightarrow$
e. $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Describing Aqueous Equations

There are three basic ways you will need to know that are used to describe aqueous chemical reactions. These are typically single and double displacement reactions. The difference between these lies in how you describe the ions of dissolved ionic compounds. Do you describe them as the separate ions, which are really what is floating around, or do you describe them by their neutral compound formula? That is, you can represent the solution as either the salt(aq) or its ions

$$NaCl(aq) \quad \text{means} \quad Na^{+}(aq) + Cl^{-}(aq) \tag{5.4.4}$$

The General (Molecular) Equation uses the formula of the salt, the total ionic equations separates soluble salts (and strong acids) into their ions, while the net ionic equation cancels any ions that appear on both the reactant and product sides of the balanced equation, the "spectator ions", as they do not participate in the reaction. The net ionic equation typically describes the actual chemistry of the reaction, and what is going on.

1. General (Molecular)Equation
2. Total Ionic Equation
3. Net Ionic Equation

In solving these problems remember, (aq) = aqueous, (s) = solid, (g) = gas and (l) = liquid, and be sure the final equation is balanced.

General (Molecular) Equation

The General or Molecular Equation describes the reactions in terms of the formula of the salts, indicating the phase. These are often (incorrectly) called "molecular equations" because they use the formula of the neutral compound as if it was a molecule. The following equation is an example of a general double displacement reaction, and given the reactants, you need to be able to predict product compounds, balance the equation and predict the phases.

$$Pb\left(ClO_4\right)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaClO_4(aq) \tag{5.4.5}$$

As indicated above, the aqueous ions separate on dissolution, and so we could write them out as separate entities, which is the Total or Complete Ionic Equation.

Total Ionic Equation

The Total or complete ionic equation writes soluble ionic compounds in terms of their ions. So if it is an ionic compound with (aq), you break it into its ions. If it is covalent, you do not break it into ions. If it is an acid, you only break it into its ions if it is strong, and that will be covered in the next section, but weak acids that are dissolved, like HF(aq) do not break into ions, while strong ones like HCl do.

Describing eq. 3.4.6 in terms of a net ionic equation gives:

$$Pb^{+2}(aq) + 2ClO_4^{-}(aq) + 2Na^{+}(aq) + SO_4^{-2}(aq) \rightarrow PbSO_4(s) + 2Na^{+}(aq) + 2ClO_4^{-}(aq) \tag{5.4.6}$$

Net Ionic Equation

The net ionic equation tells you what is happening, that is, what bonds (covalent or ionic) are being created. If you look at equation 3.4.7 you see there are two perchlorate and two sodium ions on both sides of the equation, and so we call them spectator ions because they did not react. If we cancel them the equation is still balanced and this gives the net ionic equation, which in this reaction results in the formation of the ionic compound, lead (II) sulfate.

$$Pb^{+2}\,(aq) + SO_4^{-2}\,(aq) \rightarrow PbSO_4\,(s) \tag{5.4.7}$$

This is often called a precipitate reaction as the lead (II) sulfate crashes to the bottom of the container, and some textbooks classify precipitation reactions as a class of reaction.

Example 5.4.1

Write the general, complete and net ionic equations for the reaction of **barium chloride** and **potassium sulfate**

Solution

$$\underbrace{BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KCl(aq)}_{\text{General (Molecular) Equation}}$$

$$\underbrace{Ba^{+2}(aq) + \overbrace{2Cl^-(aq) + 2K^+(aq)}^{\text{Spectator Ions}} + SO_4^{-2}(aq \rightarrow BaSO_4(s) + \overbrace{2Cl^-(aq) + 2K^+(aq)}^{\text{Spectator Ions}}}_{\text{Complete Ionic Equation}}$$

$$\underbrace{Ba^{+2}(aq) + SO_4^{-2}(aq \rightarrow BaSO_4(s)}_{\text{Net Ionic Equation}}$$

Exercise 5.4.1

Write balanced net ionic equations to describe any reaction which occur when the following solutions below are mixed.

a. $Na_2CO_3 + Sr(NO_3)_2$
b. $Cu_2SO_4 + BaCl_2$

Answer a available at the end of this section.

Answer b available at the end of this section.

Test Yourself

Homework: Section 5.4

Graded Assignment: Section 5.4

Exercise 5.4.5

Write and balance the general (molecular) equation (be sure to include phases)

a. $Al(NO_3)_3(aq) + Na_2(CO_3)(aq) \rightarrow$
b. $Pb(ClO_4)_2(aq) + Na_2SO_4(aq) \rightarrow$
c. $NaOH + CaCl_2(aq) \rightarrow$
d. $BaI_2 + H_2SO_4(aq) \rightarrow$
e. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Query 5.4.1

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Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- November Palmer, Ronia Kattoum & Emily Chaote (UALR)

Answers

NaCl is an ionic compound that dissociates into ions that conduct electricity. $C_6H_{12}O_6$ is a molecular compound that does not break up into ions, and therefore, does not conduct electricity. The sugar molecule remains intact, but each sugar molecule is separated from the other when added to water. The reason it dissolves in water is because of the term "the like dissolves the like", meaning both sugar and water are polar molecules. We will discuss this in more depth later in the text.

$Mg\left(ClO_3\right)_2 + Na_3PO_4$

$Al_2\left(SO_3\right)_3 + Na_3PO_4$

Insoluble. Sulfate ions are typically soluble, but lead(II) is an exception, so lead(II) sulfate is an insoluble solid. This means that **$PbSO_4$** (s) is not broken up into ions in water, but is a precipitate.

Soluble. perchlorate ions and alkali metals are soluble, so sodium perchlorate is soluble, meaning that it breaks up into ions, Na^+(aq) + ClO^-_4 (aq) in water.

strong electrolyte - Rule 1B

weak electrolyte - weak acid (not one of the seven strong acids)

non-electrolyte - molecular compound, not ionic

strong electrolyte- strong acid (one of the seven strong acids)

non-electrolyte - insoluble precipitate - Exception to Rule 2A

$(NH_4)_3PO_4$(aq) + **3** NaOH(aq) -> **$3NH_4OH$(aq) + Na_3PO_4(aq) no rxn**

2$(NH_4)_3PO_3$(aq) + **3**$CaSO_4$**(s)** -> **$3(NH_4)_2SO_4$(aq) + $Ca_3(PO_3)_2$ (s)**

$(NH_4)_2SO_4$(aq) + $CaCl_2$(aq) -> **$CaSO_4$(s) + 2 NH_4Cl(aq)**

$AgCH_3CO_2$ (aq) + KCl(aq) -> **AgCl(s) + KCH_3CO_2(aq)**

$Pb(NO_3)_2$(aq) + Na_2SO_4(aq) -> **$PbSO_4$(s)+ 2 $NaNO_3$(aq)**

$$Sr^{2+}(aq) + CO_3^{-2}(aq) \rightarrow SrCO_3(s)$$

$$Ba^{2+}(aq) + SO_4^{2-}(aq) -> Ba(SO_4)(s)$$

2 $Al(NO_3)_3$(aq) + **3** $Na_2(CO_3)$(aq) -> **$Al_2(CO_3)_3$(s) + 6 $NaNO_3$(aq)**

$Pb(ClO_4)_2$(aq) + Na_2SO_4(aq) -> **$PbSO_4$(s) + 2 $NaClO_4$(aq)**

2 NaOH + $CaCl_2$(aq) -> **2 NaCl(aq) + $Ca(OH)_2$**

BaI_2 + H_2SO_4(aq) -> **2 HI(aq) + $BaSO_4$(s)**

$BaCl_2$(aq) + Na_2SO_4(aq) -> **$BaSO_4$(s) + 2 NaCl(aq)**

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5.5: Neutralization Reactions

Learning Objectives

- Define the Arhennius acid/base and Bronsted-Lowry acid/base and differentiate between them
- Recognize if an acid or base is strong or weak.
- Identify the acid and base in chemical reaction
- Write molecular, complete ionic, and net ionic equations for acid-base neutralization reactions

Acids and Bases

What are acids and bases?

We will cover two definitions, the Arrhenius and Bronsted-Lowry definitions.

Arrhenius Acid/Base

Arrhenius Acid - Increases Hydronium ion to that of neutral water

Arrhenius Base - Increases hydroxide ion to that of neutral water.

This definition is based on the behavior of acids and bases when they are added to water. First we need to look at the autoionization of water. When two
water molecules bump into each other a proton can be transferred from one to the other forming hydronium and hydroxide ions (eq. 3.5.1).

$$H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq) \tag{5.5.1}$$

Arrhenius Acid

These are compounds that donate a proton to water.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq) \tag{5.5.2}$$

This is often written with the waters omitted, that is H^+(aq) represents hydronium ion and not a proton.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq) \tag{5.5.3}$$

Arrhenius Base

These are compounds that increase the hydroxide concentration. They can be soluble ionic compounds that have hydroxide ions, or compounds that remove a proton from water forming hydroxide, and the later can be molecules.

Sodium hydroxide is an ionic compound that is a strong base (you can it is soluble because of solubility rule IA)

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq) \tag{5.5.4}$$

Ammonia is a molecule that is a weak base, and it removes a proton from water forming ammonium and hydroxide. But it does not react much, so we use a two way arrow.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \tag{5.5.5}$$

In the later case the ammonia took a proton from the water and this reaction is similar to the autoinoization of water, except that it is an ammonia and not another water molecule that acquires the second proton.

NOTE: Stong Acids and Bases are strong electrolytes

Bronsted-Lowry Acid/Base

The Bronsted-Lowry definition is broader than the Arrhenius. If you look at the Arrhenius Acid, it donates a proton to water, so the Bronstead expands that to donating a proton to anything, not just water. Likewise, the Arrhenius accepts a water from water, and the Bronstead expands that to accepting a Proton from anything.

Bronstead-Lowry Acid: A Proton Donor. Note the HCl gave a proton to the water molecule in the above example

Bronstead-Lowry Base: A Proton Acceptor. Note the ammonia above accepted a proton from the water.

Extending this to a reaction that does not involve water:

$$HCl + NH_3 \rightarrow NH_4Cl \tag{5.5.6}$$

Where HCl is the acid (it donates a proton to ammonia and forms ammonium) while ammonia is the proton acceptor (becoming the ammonium ion).

YouTube Video

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Video 5.5.1: YouTube Visualization of Acid and Base behavior in water uploaded by Resa Kelly (https://youtu.be/Q7Z5jijT-Ek)

Exercise 5.5.8

Identify the following as a Bronsted-Lowry acid, base, both, or neither

a. H_3PO_4
b. $CaCl_2$
c. SO_3^{-2}
d. NaOH

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Neutralization Reactions

A reaction between an acid and a base is called a neutralization reaction, and these can be considered to be a type of displacement reaction, where the proton of the acid is being displaced as it is given to another species. You may have heard that a neutralization reaction is when an acid and a base react to form a salt plus water. That is not always true, but they do always form a salt, as the hydrogen is being "displaced" from the acid in the formation of the salt.

Neutralization reactions always give off heat, that is they are "exothermic" (a concept we will study in Chapter 5), but when a

neutralization reactions forms salt and a soluble salt, the only observation may be the heat evolved.

There are 4 types of neutralization reactions, depending on whether the acid and base are strong or weak.

1. Strong Acids and Strong Bases
2. Strong Acids and Weak Bases
3. Weak Acids and Strong Bases
4. Weak Acids and Weak Bases

To identify which type reaction is going on, you need to know if the acid or base is strong. need to know the following strong acids and strong bases. There are others, and if a problem calls it a strong acid, treat it as a strong acid. Strong acids were covered in section 3.4.1.2.1 Acids and strong bases are covered by the solubility rules Ia and IIIa of section 3.4.2.1. Table 5.5.1 reviews them for you.

Strong Acids	Strong Bases
• HCl, HBr, HI • $HClO_3$, $HClO_4$ • HNO_3 • H_2SO_4 • others as specified	• Soluble alkali metal hydroxides (solubility rule 1a) ◦ LiOH, NaOH, KOH, RbOH, CsOH • Soluble heavier alkaline earth hydroxides (solubility rule 3a) ◦ $Sr(OH)_2$, $Ba(OH)_2$ ◦ $Ca(OH)_2$ should probably be considered as a strong base even though we are not calling it a soluble salt in rule 3a. A saturated solution has is 1.8 g/L) and thus strong, and we will cover these "border cases" in gen chem 2 when we get to solubility constants.

You need to know the general (molecular), total (complete) and net ionic equations for acid base neutralization reactions

Strong Acid and Strong Base

The result is a salt plus water; the anion of the acid and the cation of the base are spectator ions. Example: HCl and NaOH

General Molecular Equation:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l) \tag{5.5.7}$$

Total Ionic Equation

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l) \tag{5.5.8}$$

Net Ionic Equation

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \tag{5.5.9}$$

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Video 5.5.1: 1'35" YouTube predicting products and writing aqueous equations for a strong acid and a strong base, (https://youtu.be/-NfFjKwCDp4). *Digital corrigendum: "all" salts of sodium (not chloride) are soluble*

Strong Acid and Weak Base

The cation of the strong acid is a spectator ion. Example: HCl & NH_3.

General Equation:

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq) \tag{5.5.10}$$

Total Ionic Equation:

$$H^+(aq) + Cl^-(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + Cl^-(aq) \tag{5.5.11}$$

Net Ionic Equation:

$$H^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq) \tag{5.5.12}$$

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Video 5.5.2: 1'06" Youtube on the reactions of strong acid and weak base (https://youtu.be/Yf2YQ6QvMEc).

Weak Acid and Strong Base

In the reaction of a Weak Acid and a Strong base the cation of the base is a spectator ion. Example: HF + NaOH.

General Equation

$$HF(aq) + NaOH(aq) \rightarrow NaF(aq) + H_2O \tag{5.5.13}$$

Total Ionic Equation

$$HF(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + F^-(aq) + H_2O(l) \tag{5.5.14}$$

Net Ionic Equation

$$HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O(l) \tag{5.5.15}$$

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Video 5.5.3: 1'19" Youtube on the reactions of weak acid and strong base (https://youtu.be/bI0tolYG_s0).

Weak Acid and Weak Base

In the reaction of a weak acid and a weak base there is no spectator ion. Since there are no spectator ions, the total ionic and the net ionic are exactly the same. Example HF and NH_3.

General Equation

$$HF(aq) + NH_3(aq) \rightarrow NH_4F(aq) \tag{5.5.16}$$

Total Ionic Equation

$$HF(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + F^-(aq) \tag{5.5.17}$$

Net Ionic Equation

$$HF(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + F^-(aq) \tag{5.5.18}$$

YouTube Video

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Video 5.5.4: 0'57" Youtube on the reactions of weak acid and strong base (https://youtu.be/561IAx1lzKk)

Contributors

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- November Palmer, Ronia Kattoum & Emily Choate (UALR)

Answers

Bronsted-Lowry acid - has 3 protons (H^+) to donate

neither - ionic compound that is charge balanced. Neither gives nor accepts protons

Bronsted-Lowry base - is an anion and so can accept a positive proton (in fact two of them).

Bronsted-Lowry base - the hydroxide ion accepts a proton and so functions as a bronstead base

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6: COUNTING MOLECULES THROUGH MEASUREMENTS



- 

6.1: Atoms and the Mole

6.2: Calculations Determining the Mass, Moles and Number of Particles

6.3: Percent Composition

6.4: Empirical Formulas

6.5: Molecular Formulas

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6.1: Atoms and the Mole

learning objectives

1. Define the mole
2. Differentiate molar mass from molecular weight
3. Calculate molar mass of a compound from its formula

The Mole

The mole is a unit of measurement that defines quantity as a counted number. It is the value of Avogadro's constant, one of the seven SI base units (review section 1B1.3.1) and the modern definition is the number of entities defined by , 6.02214076 x 10^{23} . This may not be intuitive, but this is an exact definition and there is no uncertainty to it.

1 dozen of eggs = 12 eggs

1 mole of eggs = 6.02214129 x10^{23} eggs

Note, the symbol for mole is mol and not M (which means mol/liter).

To appreciate the magnitude of Avogadro's number, consider a mole of pennies. Stacked vertically, a mole of pennies would be 4.5 × 10^{17} mi high, or almost six times the diameter of the Milky Way galaxy. If a mole of pennies were distributed equally among the entire population on Earth, each person would have more than one trillion dollars. The mole is so large that it is useful only for measuring very small objects, such as atoms.

Students need to memorize this number to 4 significant digits (6.022 x 10^{23}) but may need more if they use it in a calculation where using 4 significant digits determines the wrong number of significant digits in the answer. So if you multiply a six digit number like 1.00000 by Avogadro's number, you should use 7 digits for avagadro's number, ie., 6.022140 x 10^{23} .and the answer is 6.02214 x 10^{23} and not 6.022 x 10^{23}.

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Video 6.1.1 Ted-Ed video on "How big is a mol?" by Daniel Dulek

Molar Mass (Formula Weight)

The molar mass (formula weight) is the mass of 6.022x10^{23}) atoms or compounds (molecules or ionic formula). The terms Molar Mass and Formula Weight are synonymous. Molecular weight is also often used to describe molar mass, but formula weight is better, as not all compounds are molecules (many are ionic).

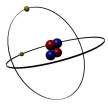
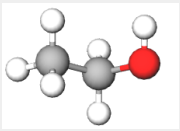
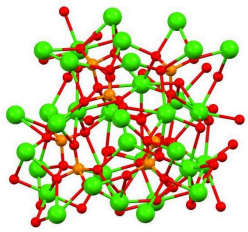
Substance (formula)	Atomic, Molecular, or Formula Mass (u)	Molar Mass (g/mol)	Structure
Helium (He)	4.0026 (atomic mass)	4.0026	 Pixabay
ethanol (C_2H_5OH)	46.069 (molecular mass)	46.069	
calcium phosphate $[Ca_3(PO_4)_2]$	310.177 (formula mass)	310.177	 wikimedia commons

Figure 6.1.1: Comparison of atomic and molar mass for atom, molecule and ionic compound

The molar mass is often called the molecular weight or the formula weight (*fw*). Technically speaking, molecular weight only works for molecules, where it is the weight of one mole of molecules, where each molecule is a particle. For ionic compounds you could have one piece of crystal that has a mole of its constituent elements as defined by the ionic formula, but there is only one particle, and so it is better to consider its formula mass, or formula weight. We will use the symbol fw to indicate formula weight.

Molar Mass of Elements

The importance of a mole is that if you have one mole of an element, the mass of the mole in grams is the same as the average mass of each element in Daltons (u). So if the average isotopic mass of an atom in a sample of carbon weights 12.011 u, then one mole of a sample of carbon weights 12.011 grams. Note, there is not a single atom of copper that weighs 63.546 amu, but a mole of coppers does weigh 63.546 grams. So the atomic weight on the periodic table is the molar mass, which is the average mass of 6.022×10^{23} atoms of that element in units of grams. It is an equivalence statement.

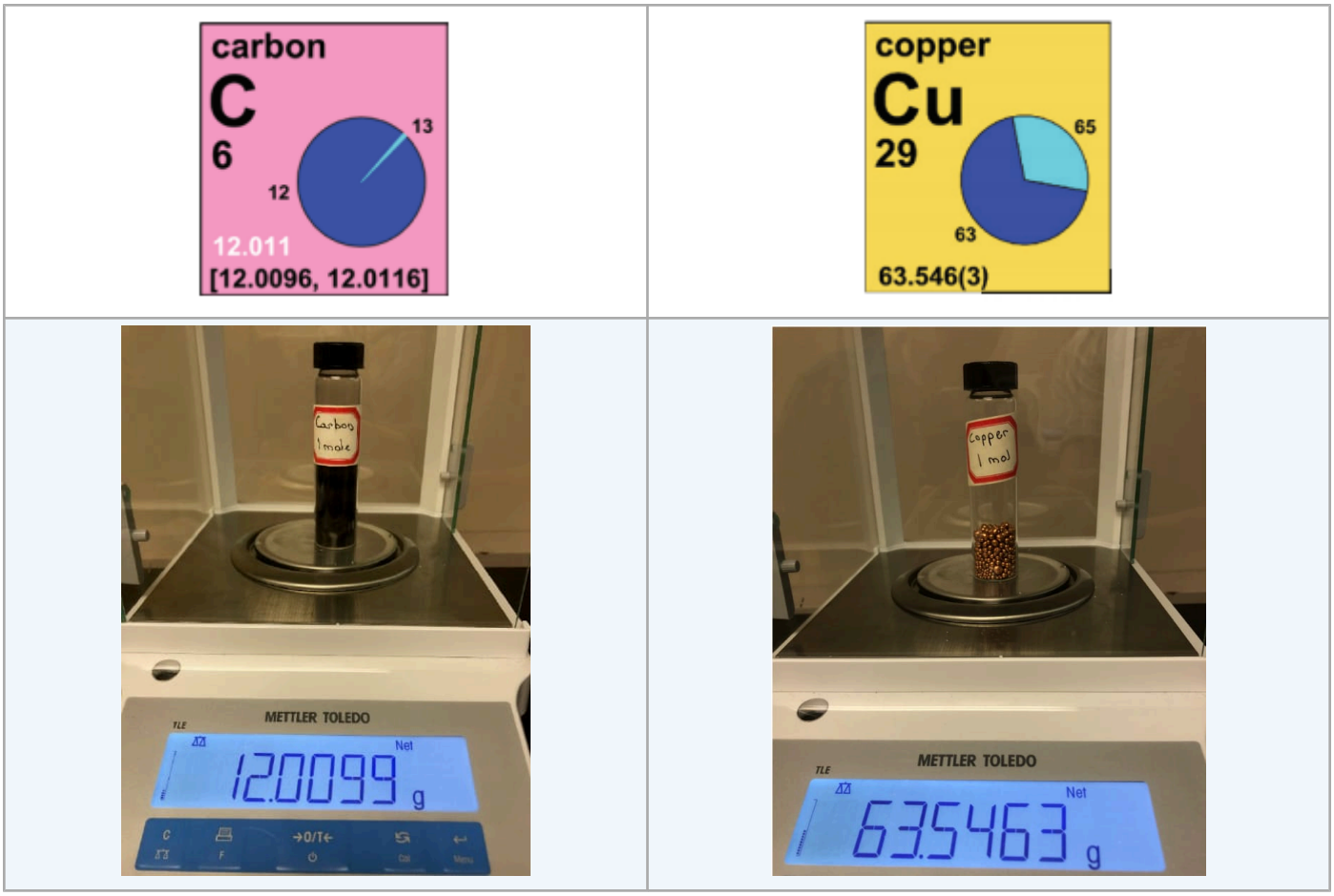


Figure 6.1.2: The two samples of carbon (left) and copper (right) have the same number of molecules (1 mol or 6.022×10^{23}) but different masses and volumes. Note, these are tared weights, that is, the weight of the container was subtracted from the displayed value (see use of analytical balances).

Exercise 6.1.1

Which has a greater number of particles? Which of the following has a greater mass?

A) 2.5 moles of He

B) 2.5 moles of Ar

Answer available at the end of this section.

Molar Mass of Compounds

Molecular

Ibuprofen, $C_{13}H_{18}O_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	13	×	12.01	=	156.13
H	18	×	1.008	=	18.114
O	2	×	16.00	=	32.00
				Molecular mass	206.27

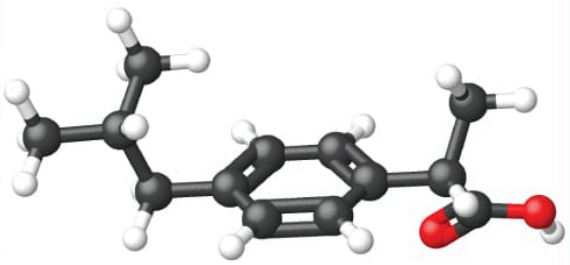


Figure 6.1.3: Calculation of molar mass of molecule (ibuprofen) (adopted from OpenSTAX)

Ionic

Remember Ionic compounds are discrete cations and anions combined to form a neutral crystal lattice. The formula for an ionic compound is the lowest whole number ratio of cations to anions and is not defined by the bonds, and there is no discrete "molecule" like there was for ibuprofen. Sodium chloride for example has the formula of NaCl, even though each sodium is bonded to 6 chlorines, while each chlorine in turn if bonded to 6 sodiums forming a lattice in 3 diensional space. Sodium chloride is an ionic compound composed of sodium cations, Na^+, and chloride anions, Cl^-, combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (Figure 6.1.4), even though a real sample has many more sodium and chloride ions within its crystal structure.

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

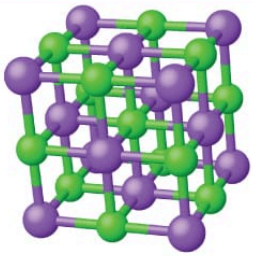


Figure 6.1.4: Calculation of molar mass of ionic compound (table salt, NaCl) (adopted from OpenSTAX)

You can also calculate the molar mass of polyatomic ions. This is often called molecular weight or formula weight.

Example: The molar mass of sulfate (SO_4^{-2})

1 mole S (1)(32.07) = 32.06g
4 mole O (4)(16.00) = 64.00g
1 mole SO_4^{-2} = 96.06g

Note

The molar mass is an **Equivalence Statement** and like any equivalence statement, it results in two conversion factors (number to mass, or mass to number).

For Sodium Chloride the formula weight is 58.44g/mol

$$58.44\ g\ NaCl = 1\ mol\ NaCl \tag{6.1.1}$$

To convert grams to moles you divide by the molar mass

$$\left(\frac{1\ mol\ NaCl}{58.44\ g\ NaCl}\right) \tag{6.1.2}$$

To convert from moles to grams you multiply by the molar mass

$$\left(\frac{58.44\ g\ NaCl}{1\ mol\ NaCl}\right) \tag{6.1.3}$$

Test Yourself

Homework: Section 6.1

Exercise 6.1.2

What is the molar mass for each of the following molecules?

a. Aluminum Sulfate
b. Benzene (C_6H_6)
c. Acetylene (C_2H_2)
d. Water
e. Hydrogen Peroxide

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Exercise 6.1.3

What is the molar mass for each of the following molecules?

a. Sodium Bicarbonate
b. Ammonium Nitrate
c. Potassium Perchlorate
d. Silver Chloride
e. Acetic Acid

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Contributors and Attributions

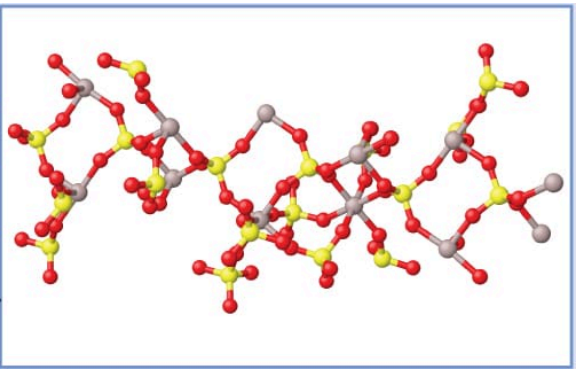
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- November Palmer, Ronia Kattoum & Emily Choate (UALR)
- Some material adapted from Paul Flowers, Klaus Theopold & Richard Langley via OpenStax (LibreText Copy section 3.1)

Answers

Both have the same number of particles because there are 2.5 moles of each. That is similar to saying 2.5 dozen of donuts is the same number of items as the 2.5 dozen of cars. However, 2.5 moles of cars is going to have a larger mass than 2.5 dozen of donuts. By the same token, 2.5 moles of Ar is going to have more mass than 2.5 moles of He, even if they have the same number of particles.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
				Molecular mass	342.14



Aluminum Sulfate ($Al_2(SO_4)_3$) = **342.14 g/mol**(Image adopted from OpenSTAX)

NOTE: we could have used the molar mass of sulfate that was calculated above

$$2(Al) + 3(\text{Sulfate})$$
$$2(26.98) + 3(96.06) = 342.14$$

Benzene (C_6H_6) = **78.12g/mol**

Acetylene (C_2H_2) = **26.04g/mol**

Water (H_2O) = **18.02g/mol**

Hydrogen Peroxide (H_2O_2) = **34.02g/mol**

Sodium Bicarbonate ($NaHCO_3$) = **84.02g/mol**

Ammonium Nitrate (NH_4NO_3) = **80.08g/mol**

Potassium Perchlorate ($ClKO_4$) = **138.54g/mol**

Silver Chloride (AgCl) = **143.32g/mol**

Acetic Acid (CH_3COOH) = **60.06g/mol**

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6.2: Calculations Determining the Mass, Moles and Number of Particles

learning objectives

- Convert grams of a substance to moles
- Convert moles of a substance to grams

Counting by Measuring

In this section we will learn to "count" the number of particles of a solid substance by measuring its mass. This is building off of section 2.8.2.2, molar mass of a compound and Dimensional analysis (section1B.4), where the molar mass is considered an equivalent statement that allows you to convert from a measured value of mass to a counted value of number of particles. As such, it is like a unit conversion based on a physical constant (section 2B.3.5.3), the molar mass of a compound.

We will use the following symbols

m = mass (g)

n = moles (mol)

Formula Weight is defined as

$$fw = \frac{m(g)}{n(mol)} \tag{6.2.1}$$

$$mass = moles\ times\ formula\ weight$$

This formula weight is an equivalence statement that depends on the identity of the chemical and converts between its mass and its number of moles. For water

$$1\ mol H_2O = 18\ grams \tag{6.2.2}$$

$$fw(H_2O) = 18\frac{g}{mol}$$

$$m = n(fw) \tag{6.2.3}$$

and there are thus two conversion factors

$$\underbrace{\frac{m}{fw}}_{\text{converts mass to mols}} \quad \leftarrow or \rightarrow \quad \underbrace{n(fw)}_{\text{converts mols to mass}} \tag{6.2.4}$$

That is, dividing the mass by the formula weight converts it to the number of moles, and multiply the molese by the formula weight converts its moles to mass. The formula weight relates mass to moles the way density relates mass to volume, and it is a constant that for any element or compound that can be determined from the periodic table.

Grams to Moles

Figure $6.2.1$ shows the flow chart for converting mass to moles. I

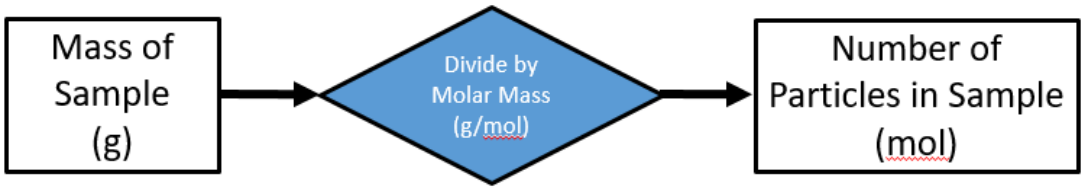


Figure 6.2.1: Flowchart for converting mass to moles.

The best way to solve these problems is to set up a conversion factor based on the equivalence statement (Equation 2.9.2), and then multiply by a conversion factor based on this equivalence statement, where you put the mass in the denominator (which is the same as dividing by the molar mass).

When Using Conversion Factors Keep Units in Equations

Example 6.2.1

How many molecules are in 3.00g of ibuprofen?

Solution

We calculated the molar mass of ibuprofen in section 2.8.2.2.1 to be 206.27 g/mol.

From the molar mass we can generate the equivalence statement:

$$206.27g\ C_{13}H_{18}O_2 = 1\ mole\ C_{13}H_{18}O_2 \tag{6.2.5}$$

We are converting a mass of 3.00 g, which has 3 significant Figures, and so will use one extra for the molar mass, and use 206.3 g/mol. From

$$3.00\,\cancel{g}\ C_{13}H_{18}O_2 \underbrace{\left(\frac{mol\ C_{13}H_{18}O_2}{206.3\ \cancel{g}}\right)}_{\text{converts mass to mol}} = 0.0145 \text{ mol or } 14.5\text{mmol} \tag{6.2.6}$$

or we can convert to actual molecules, but need a second equivalence statement

mole = 6.02214x10^{23} molecules ibuprofen

$$3.00\,\cancel{g}\ C_{13}H_{18}O_2 \underbrace{\left(\frac{\cancel{mol}\ C_{13}H_{18}O_2}{206.3\ \cancel{g}}\right)}_{\text{converts mass to mol}} \underbrace{\left(\frac{6.022x10^{23}\ C_{13}H_{18}O_2}{\cancel{mol}}\right)}_{\text{converts moles to molecules}} = 8.76x10^{21} \text{ molecules of ibuprofen} \tag{6.2.7}$$

NOTE: Either Answer is Correct.

Moles to Grams

Figure 6.2.2: shows the flow chart of converting number of particles to mass, but once again, we will use the conversion factor based on the equivalence statement.



Figure 6.2.2: Flow chart for converting number of particles to mass.

The Trick: The Molar Mass is a conversion factor like a Physical Constant, and so make a Equivalence Statement out of it.

Example 6.2.2

Using the molar mass of 206.27 g/mol, calculate the mass in grams of 10 molecules of ibuprofen?

Solution

To answer this we need two equivalence statements

- **1 mole = 206.27 g ibuprofen**
- **1 mole = 6.02214x10^{23} molecules ibuprofen**

Note: 10 molecules is an exact number and has no significant digits Since the molar mass was given with five significant digits, we would not use 6.022x10^{23} for Avogadro's number because that has 4, and so we have to decide how many to use with Avogadro's number. The rule of thumb (see below) when deciding how many to use is to use one more significant digit than the number you are working with, and so we choose six.

$$10\ \cancel{molecules}\ C_{13}H_{18}O_2 \underbrace{\left(\frac{mol\ C_{13}H_{18}O_2}{6.02214\ x10^{23}\ \cancel{molecules}}\right)}_{\text{converts molecules to moles}} \underbrace{\left(\frac{206.27g\ C_{13}H_{18}O_2}{\cancel{mol}}\right)}_{\text{converts moles to mass}} = 3.4252x10^{-21}\ \text{g of ibuprofen} \qquad (6.2.8)$$

Exercise 6.2.1

Answer the following Questions

a. How many moles is 12.2 g of Sodium Bicarbonate?
b. How many moles is 18.8 g of Ammonium Nitrate?
c. How many atoms is 30.8 g of Ammoium Nitrate?
d. What is the mass of 2.45 mole Silver Chloride?
e. what is the mass of 3.44 mole of benzene (C_6H_6) ?

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Significant Digits in Molar Mass Calculations

The number of significant digits you use in calculating the molar mass depends on your data. The rule of thumb is use one more significant digit than the measured value that determines the number of significant digits in your final answer. There is no need to use more, but if you use less, then the significant digits of the molar mass determines the significant digits of your final answer.

You should redo equation 2.9.5 using a molar mass of 206.27 g/mol and you will see the final answer (to three significant digits) is the same. In making these conversions, you should always use a molar mass with one extra significant digit than your measured number. If you had uses 3.00000g of ibuprofen (6 significant digits), then if possible you should have used seven significant digits for the molar mass. If you had used five significant digits (206.27g/mol) your answer would have 5 significant digits, even though your measured value had seven.

Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is

the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- some material adapted from Paul Flowers, Klaus Theopold & Richard Langley via OpenStax (LibreText Copy section 3.1)

Answers

Sodium Bicarbonate ($NaHCO_3$) = **0.145 mol**

Ammonium Nitrate (NH_4NO_3) = **0.235 mol**

Potassium Perchlorate ($ClKO_4$) = **0.222mol**

Silver Chloride (AgCl) = **351g**

Benzene (C_6H_6) = 269**g**

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6.3: Percent Composition

learning objectives

1. Introduce Mass Percent and Mole Fraction
2. Calculate Mass Percent from Chemical Formulas

Introduction

Percent and fraction essentially tell you the same thing, they describe the ratio of a part of the whole. Because a chemical compound has a constant composition, which is defined by is molecular or compound formula, the fraction of each type of element must be constant. For example, water has 3 atoms, one hydrogen and two oxygens. So the fraction of oxygen is 1/3 (33.3%) and the fraction of hydrogen is 2/3rd (66.7%), as exemplified in the left pie chart of Figure 6.3.1. But water weights 16 times as much as hydrogen and so if you could weigh a water molecule, the mass of oxygen would be 16 Daltons, while that of the two hydrogens would be 2 Daltons, so on a mass basis, water if 88.9% oxygen and 11.1% hydrogen. So when we say the fractional or percent composition, we need to identify how we are determining it. Usually, but not always, chemists use percentage for mass and fraction for moles (number of particles).

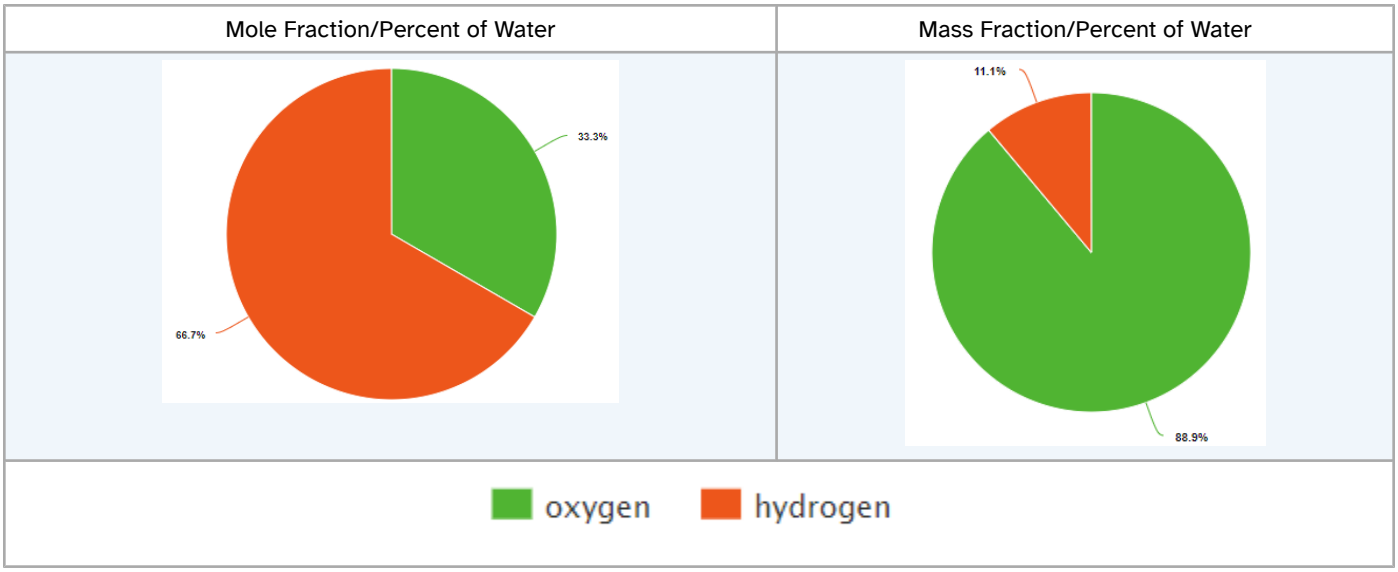


Figure 6.3.1: Mole (left) and mass (right) fractions (or percentages) of oxygen and hydrogen in water.

So, is water mostly oxygen or hydrogen? Well, that depends on how you look at it. If you are counting atoms, it is mostly hydrogen. If you are measuring the mass, it is mostly oxygen.

Mass Percent

Mass percent of any element in a compound is 100 time the mass of the element divided by the total mass

$$\text{Mass \% element} = \frac{\text{mass element}}{\text{mass compound}} \times 100\% \tag{6.3.1}$$

One mole of water weights 18 grams, with 2 grams being hydrogen and 16 being oxygen, so the fraction of hydrogen is 2/18 (0.1111) and the fraction oxygen is 16/18 (.8889), and note that the sum of the fractions is 1. The percentage of an element in a compound is 100 times the fraction, so for water the mass percent hydrogen is 11.11% and the mass percent oxygen is 88.89%.

$$\text{Mass \%H} = \frac{\text{mass H}}{\text{mass compound}} \times 100\% \tag{6.3.2}$$

$$\text{Mass \%O} = \frac{\text{mass O}}{\text{mass compound}} \times 100\% \tag{6.3.3}$$

for water

$$\text{Mass \%H} = \frac{2\text{g}}{18\text{g}} \times 100\% = 11.11 \tag{6.3.4}$$

$$\text{Mass \%O} = \frac{16\text{g}}{18\text{g}} \times 100\% = 89.89 \tag{6.3.5}$$

Equations 2.10.4 and 2.10.5 describe the right pie chart in Figure 6.3.1, and note they add to 100%.

Video 6.3.1 shows a technique to calculate the percent composition of a formula that minimizes the amount of work you need to do.

YouTube Video

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View this YouTube Video online

Video 6.3.1: Calculating the mass % composition of the elements in acetic acid.

NOTE: Even though mass has units, mass percent is unit-less, and so care needs to be taken to identify what the percent or fraction relates to. Also, because a sample of a pure substance consist of molecules or ionic compounds of the same formula, the percent or fractional value does not depend on the sample size. So one molecule of water has the same percent composition as a gallon of [pure] water.

It should also be noted that as the following two videos show, some formulas have the same percent composition, even though they are different molecules, like acetylene (C_2H_2) and benzene (C_6H_6).

YouTube Video

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View this YouTube Video online

Video 6.3.2: Comparing % Composition of acetylene with benzene.

Exercise 6.3.1

What is the mass percent of oxygen in dichlorine heptoxide?

Answer available at the end of this section.

Mole Fraction

The mole fraction of each element in a sample is the number of moles of that in a sample divided by the total number of moles in the sample, and this can be determined by the chemical formula, where it is the number of atoms of that element in the formula divided by the total number of atoms in the formula.

$$\text{mole fraction of an element } = \frac{\text{moles of the element}}{\text{total moles in compound}} \tag{6.3.6}$$

shows the numbers of each type of element and so one mole of water (H_2O) has two moles of hydrogen and one mole of water. That is the mole fraction of hydrogen is 2/3 (.6667) and the mole fraction oxygen is 1/3 (.3333), meaning the mole fraction for water is 67% hydrogen and 33% oxygen.

$$\text{mole fraction H} = \frac{\text{mole H}}{\text{total mol in compound}} \tag{6.3.7}$$

$$\text{mole fraction O} = \frac{\text{mole O}}{\text{total mol in compound}} \tag{6.3.8}$$

for water

$$\text{mole fraction H} = \frac{2 \text{ mole H}}{3 \text{ mol}} = 0.667 \tag{6.3.9}$$

$$\text{mole fraction O} = \frac{1 \text{ mole O}}{3 \text{ mol}} = 0.333 \tag{6.3.10}$$

Equations 2.10.9 and 2.10.10 describe the right pie chart in Figure 6.3.1.

Exercise 6.3.2

What is the mole fraction of oxygen in dichlorine heptoxide?

Answer available at the end of this section.

Molecular Composition in Calculations

Exercise 6.3.3

How many moles of oxygen are in 12 moles of dichlorine heptoxide?

Answer available at the end of this section.

Exercise 6.3.4

How many grams of oxygen are in 1200 mg of dichlorine heptoxide?

Answer available at the end of this section.

Test Yourself

Query 6.3.1

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- Ronia Kattoum (Learning Objectives)
- Liliane Poirot

Answers

Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g

Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g

Molar mass of Cl_2O_7 = 182.90 g/mol

$$\% \text{ O} = \frac{112.00 \text{ g O}}{182.90 \text{ g}} \times 100\% = 61.24\% \text{ O}$$

Moles of Cl in 1 mol Cl_2O_7 = 2 mol Cl

Moles of O in 1 mol Cl_2O_7 = 7 mol

Total moles in Cl_2O_7 = 9

$$mole\ fraction \text{ O} = \frac{7 \text{ mol O}}{9 \text{ total moles}} = .778$$

Moles of O in 1 mol Cl_2O_7= 7 mol

$$12\ mol\ Cl_2O_7 \times \frac{7 mol O}{1\ mol\ Cl_2O_7} = 84\ mol\ O$$

Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g

Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g

Molar mass of Cl_2O_7 = 182.90 g/mol

$$1200\ mg\ Cl_2O_7 \times \frac{1\ g}{1000\ mg} \times \frac{112.00\ g\ O}{182.90\ g\ Cl_2O_7} = 0.7348\ g\ O$$

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6.4: Empirical Formulas

learning objectives

1. Identify the difference between empirical and molecular formulas
2. Determine empirical formula formula using percent composition
3. Determine empirical formula using mass data

Introduction

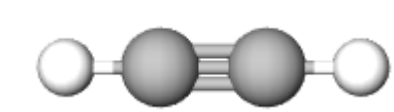
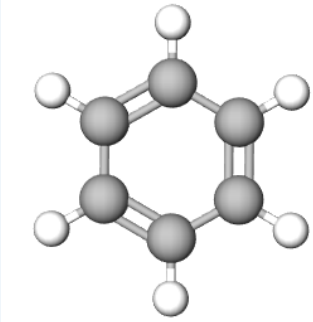
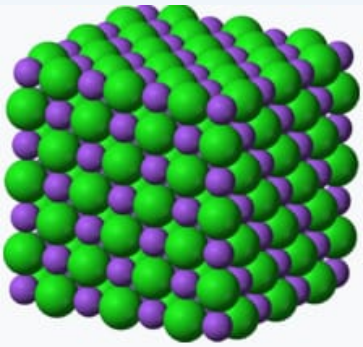
Empirical measurements are based on a measurable (empirical) quantity like mass. Knowing the mass of each element in a compound we can determine its formula. There are two types of formulas, empirical and molecular.

Empirical Formula: Lowest whole number ratio of the elements in a compound

Molecular Formula: Actual whole number ratio of the elements in a compound. We will go over these in the next section

The Empirical formula is the lowest whole number ratio of the elements in a compound. In video 6.3.2 (section 6.3), we discovered that benzene and acetylene have the same mass percent composition, and thus it is logical that they have the same ratio of elements to each other, that is, they have the same empirical formula.

Table 6.4.1 *Empirical and Molecular Formula of some compounds*

Compound	Acetylene	benzene	NaCl
Structure			
Empirical Formula	CH	CH	NaCl
Molecular Formula	C_2H_2	C_6H_6	not a molecule (use NaCl)

Empirical Formula

The Empirical formula is the lowest whole number ratio of the elements in a compound. In (section 2.10), we discovered that benzene and acetylene have the same mass percent composition, and thus it is logical that they have the same ratio of elements to each other, that is, they have the same empirical formula.

For salts that do not have homonuclear diatomic ions (like Hg_2^{+2} or O_2^{-2}) the empirical formula is the formula we write to describe the salt. Mercury(I)chloride has the empirical formula of HgCl, but the real compound formula is Hg_2Cl_2 (review table 2.7.3)

Multiple molecules can have the same empirical formula. For example, benzene (C_6H_6) and acetylene (C_2H_2) both of the empirical formula of CH (see Figure 6.4.1.

Calculating Empirical Formulas

Steps:

1. **Obtain Mass of Each Element (in grams)**
   - if given total mass you need the mass of all but one element (sum of masses equals total mass).
   - if given % composition assume 100 g and convert to mass.
     - if given % composition you need data for of all but one element (sum of percents equals 100%).
2. **Calculate # of moles of each element present** (from masses and atomic weights).
   - You now have a formula representing the mole ratio of the elements in the compound and you need to make these integers.
3. **Divide # of Moles of each element by the one with the smallest value** (forcing it to one and making all other values greater than one).
4. **Multiply the results of step 3 by the smallest integer which will convert them all to whole numbers.**

Note

The trick is to convert decimals to fractions and then multiply by the lowest common denominator (watch video 6.4.1)

0.5=$\frac{1}{2}$	0.33=$\frac{1}{3}$	0.67=$\frac{2}{3}$	0.25=$\frac{1}{4}$	0.75=$\frac{3}{4}$
0.2=$\frac{1}{5}$	0.4=$\frac{2}{5}$	0.6=$\frac{3}{5}$	0.8=$\frac{4}{5}$	0.167=$\frac{1}{6}$

The following video shows how to calculate the empirical formula for aspiring.

Calculate the empirical formula for aspirin:

Aspirin is made of H, O & C, and was analyzed to contain 60.0% carbon and 35.5% Oxygen.

YouTube Video

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Video 6.4.1: Empirical formula of aspirin

Example 6.4.1

A certain compound was found to contain 67.6% C, 22.5% O, and 9.9% H. What is the empirical formula?.

Solution

$$22.5gO\left(\frac{1molO}{16.00g}\right)=1.4\Rightarrow\frac{1.4}{1.4}=1 \tag{6.4.1}$$

$$67.6gC\left(\frac{1molC}{12.011g}\right)=5.63\Rightarrow\frac{5.63}{1.4}=4 \tag{6.4.2}$$

$$9.9gH\left(\frac{1molH}{1.007g}\right)=9.9\Rightarrow\frac{9.9}{1.4}=7 \tag{6.4.3}$$

C_4H_7O

This Applet comes from the ChemCollective at Carnegie Mellon University. This link will send you to the video and tutorial associated with this applet. Every time you load the page a new problem will load, and there are a series of tiered hints to help you work through the problems.

Test Yourself

Homework: Section 6.4

Graded Assignment: Section 6.4

Exercise 6.4.1: empirical formula

Calculate the Empirical formula for the following

a. A 3.3700 g sample of a salt which contains copper, nitrogen and oxygen, was analyzed to contain 1.1418 g of copper and 1.7248 g of oxygen.
b. A compound of nitrogen and oxygen that contains 30.43% N by weight.
c. A 2.402-g sample of made of C, H, N and O contains 1.121 g of N, 0.161 g H, 0.480 g C, and an unspecified amount of oxygen.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 6.4.2: empirical formula

Given the empirical formula of the compound in part (a) of the above:

a. What could be the formula of the compound.
b. What would be the name of the compound, using the formula in the answer from part (a) of this question.

Answer a available at the end of this section.

Exercise 6.4.3: empirical formula

In section 2.10.2 we saw that benzene and acetylene both have the same mass percent composition (92.3% C and and 7.7% H), so calculate their empirical formulas

a. What is the empirical formula of acetylene (C_2H_2)?
b. What is the empirical formula of benzene (C_6H_6)? (Knowing (a) the empirical formula of acetylene and (b) that benzene and acetylene have the same mass % composition, cam you answer this without doing any math?)

Answer a available at the end of this section.

Contributors and Attributions

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- November Palmer & Emily Choate (UALR)

Answers

CuN_2O_6

NO_2

N_2H_4CO

$Cu(NO_3)_2$

Copper(II) nitrate of cupric nitrate

CH

CH

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6.5: Molecular Formulas

learning objectives

1. Distinguish between empirical formula and molecular formula
2. Determine empirical formula and molecular formula using percent composition
3. Determine empirical formula and molecular formula using mass data

Introduction

Empirical measurements are based on a measurable (empirical) quantity like mass. Knowing the mass of each element in a compound we can determine its formula. There are two types of formulas, empirical and molecular.

Empirical Formula: Lowest whole number ratio of the elements in a compound

Molecular Formula: Actual whole number ratio of the elements in a compound.

The Empirical formula is the lowest whole number ratio of the elements in a compound. In (section 2.10), we discovered that benzene and acetylene have the same mass percent composition, and thus it is logical that they have the same ratio of elements to each other, that is, they have the same empirical formula.

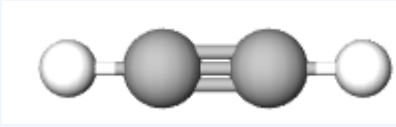
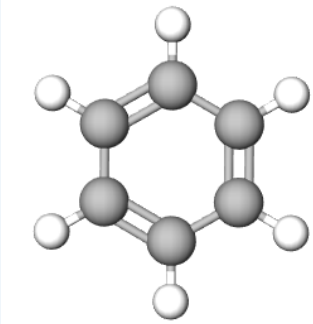
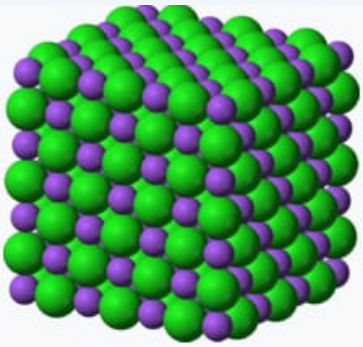
Compound	Acetylene	benzene	NaCl
Structure			
Empirical Formula	CH	CH	NaCl
Molecular Formula	C_2H_2	C_6H_6	not a molecule (use NaCl)

Figure 6.5.1: Empirical and molecular formulas of several simple compounds.

Molecular Formula

To calculate the molecular formula we need additional information beyond that of the mass or mass percent composition, we need to know the molar mass of the substance. There are many experimental ways that can be determined, and we will learn some as the semester proceeds. The mass spectrometer that we used to determine the isotopic composition in section 2.3 of this Chapter could be used to determine the molar mass of many unknowns. But there are other techniques, and at this point in the semester, the molar mass will be treated as a given.

The empirical formula represents the lowest whole number ratio of the elements in a molecule while the molecular formula represents the actual formula of the molecule. So the molecular formula must be an integer multiple of the empircal formula, that is it is n times larger where n=1 or 2 or 3 or.... but it is an integer. This gives the following relationship

$$[\text{Molecular Formula} = \text{n}([\text{Empirical Formula})] \tag{6.5.1}$$

then if we have one mole

$$[\text{Molecular Weight} = \text{n}([\text{Empirical Weight})] \tag{6.5.2}$$

or

$$n = \frac{[\text{Molecular Weight}]}{[\text{Empirical Weight}]} \tag{6.5.3}$$

So you calculate the Empirical formula as above, then determine the weight of one mole, divide that into the molar mass, and that tells you how many times it is bigger, and then multiple the emprical formula by that number.

Let's compare Benzene to acetylene. In video 2.10.2 (section 2.10.2) we saw that benzene and acetylene had the same mass % composition and in exercise 2.11.3 we say they had the same empirical formula

Table 6.5.1: data setup for calculating molecular formula for empirical formula and molar mass for benzene and acetylene

name:	Benzene	Acetylene
Formula Weight	78.12g/mol	26.04g/mol
Empirical Formula	CH	CH
Empirical Formula Weight	13.017	13.017
n=$\frac{[\text{Molecular Weight}]}{[\text{Empirical Weight}]}$	$\frac{78.12}{13.017} = 6$	$\frac{26.04}{13.017} = 2$
Formula = n times empirical formula	6(CH) = C_6H_6	2(CH) = C_2H_2

Calculating Molecular Formulas

Example 6.5.2

A certain compound was found to contain 67.6% C, 22.5% O, and 9.9% H. If the molecular weight of the compound was found to be approximately 142 g/mol, what is the correct molecular formula for the compound?

Solution

C_4H_7O has a empirical weight of

(4)(12.01) + (7)(1.007) + (16.00) = 71.09g/mol

$$\left(\frac{142g/mol}{71.09g/mol}\right) = 2 \tag{6.5.4}$$

multiply the coefficients by 2

$C_8H_{14}O_2$

Test Yourself

Homework: Section 6.5

Graded Assignment: Section 6.5

Exercise 6.5.4: Molecular formula

Calculate the molecular formula for the following

a. A compound has an empirical formula of C_2HF has a molar mass of 132.06 g/mol.
b. 200.0 g sample of an acid with a molar mass of 616.73g/mol contains 171.36 g of carbon, 18.18g of nitrogen and the rest is hydrogen.
c. Strychine has a molar mass of 334 g/mol and percent composition of 75.42%C, 6.63%H and 8.38%N and the rest oxygen.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Contributors and Attributions

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- November Palmer & Emily Choate (UALR)
- Ronia Kattoum (Learning Objectives)

Answers

$C_6H_3F_3$

$C_{44}N_4H_{32}$

$C_{21}N_2H_{22}O_2$

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7: STOICHIOMETRY

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- Topic hierarchy

7.1: Stoichiometric Reactions

7.2: Theoretical Yield, Limiting and Excess Reagents

7.3: Percent Yield

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7.1: Stoichiometric Reactions

Learning Objective

- Master Stoichiometric Relationships
  - Determine relationships of quantities of reactants consumed and products produced
  - express these in both number of chemical entities (moles) or masses of chemical entities

Introduction

A balanced chemical equation gives the identity of the reactants and the products as well as the accurate number of molecules or moles of each that are consumed or produced in a chemical reaction. **Stoichiometry** is a collective term for the quantitative relationships between the masses, the numbers of moles, and the numbers of particles (atoms, molecules, and ions) of the reactants and the products in a balanced chemical equation. The stoichiometric coefficient (section 5.1.3) represents the lowest whole number of each reactant or product consumed or produced as related to each other and the ratio of the coefficients represents the relative rates at which the species are consumed or produced. This, for each specific reaction, the coefficients represent an equivalence statement. Consider the following balanced equation describing the formation of Analine ($C_6H_5NH_2$) and Magnetite (Fe_3O_4 or Iron (II,III)oxide) from nitrobenzene ($C_6H_5NO_2$), iron and water. [Note: you will not be required to name compounds like Fe(II,III)oxide, where the cation can have more than one oxidation state, but it is common. Here one iron lost 2 electrons and the other 3, and they form part of a ionic crystal lattice with oxide being the counter ion).

$$4C_6H_5NO_2 + 9Fe + 4H_2O \rightarrow 4C_6H_5NH_2 + 3Fe_3O_4 \tag{7.1.1}$$

The **balanced equation results in a set of Equivalence Statements for a specific chemical reaction** that relate the quantity of reactants and products that are consumed or produced to each other. This equivalence statement is only valid for the equation being described. For the above equation we can derive the following equivalence statements

$4mol\ C_6H_5NO_2\ consumed\ =\ 9mol\ iron\ consumed$	$9mol\ Fe\ consumed\ =\ 4mol\ C_6H_5NH_2\ produced$
$4mol\ C_6H_5NO_2\ consumed\ = 4mol\ H_2O\ consumed$	$9mol\ Fe\ consumed\ =\ 3mol\ Fe_3O_4\ produced$
$4mol\ C_6H_5NO_2\ consumed\ =\ 4mol\ C_6H_5NH_2\ produced$	$4mol\ H_2O\ consumed\ =\ 4mol\ C_6H_5NH_2\ produced$
$4mol\ C_6H_5NO_2\ consumed\ =\ 3mol\ Fe_3O_4\ produced$	$4mol\ H_2O\ consumed\ =\ 3mol\ Fe_3O_4\ produced$
$9mol\ Fe\ consumed\ =\ 4mol\ H_2O\ consumed$	4 $mol\ C_6H_5NH_2\ produced\ =\ 3mol\ Fe_3O_4\ produced$

It needs to be noted that each of the above equivalent statements can produce two conversion factors, and that the "entity" in the numerator is different than the entity in the denominator, and thus is is important that you include the entity when you do calculations. If we look at the first equivalence statement,

$$4mol\ C_6H_5NO_2\ consumed\ =\ 9mol\ iron\ consumed \tag{7.1.2}$$

we get the following two conversion factors

$$\left(\frac{4mol\ C_6H_5NO_2}{9mol\ iron}\right)\ \ or\ \ \left(\frac{9mol\ iron}{4mol\ C_6H_5NO_2}\right) \tag{7.1.3}$$

The first converts moles iron to moles nitrobenzene, and the second converts moles nitrobenzene to moles iron.

Counting by Measurement

So how do we "count" chemical entities involved in a chemical reaction?

We count chemical entities in a reaction by measuring macroscopic observables like mass, volume, temperature and pressure. What we measure depends on the state of matter, but we need to relate a measurable property to the number of particles. In this

Chapter we will learn how to do this for solid substances and for solutes in a solution. In the gas phase Chapter we will use the ideal gas law (PV=nRT) to do this for gas phase reactants.

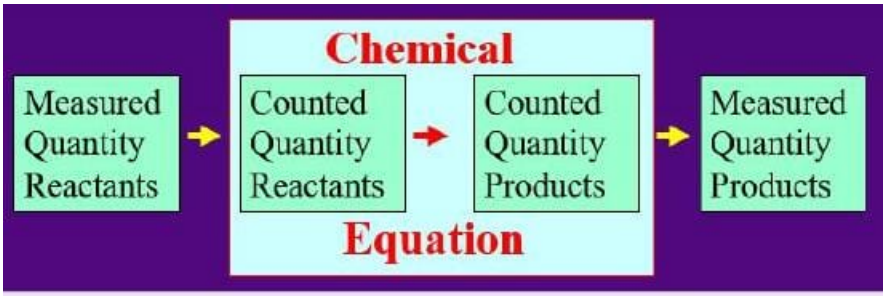


Figure 7.1.1: Flow chart showing three steps for using a balanced chemical equation to relate measured quantities of various reactants and products to each other.

We will start with the simplest types of stoichiometric equations, those involving masses.

Calculations Involving Mass Measurements

Mass measurements are the easiest as you use the molar mass to convert between grams and moles. In Figure 7.1.1 you see three arrows, two yellow and one red. Each of these involves a conversion factor. The yellow ones are conversions involving the molar masses, which relate the mass (grams) of a substance to its moles, and the red one involves the balanced chemical equation, which relates the stoichiometric coefficients of the various species to each other.

Technique: This is a three step process which should be done in one equation which uses three conversion factors.

- Conversion Factor #1: Use molar mass to convert **mass** of known material **to moles**.
- Conversion Factor #2: Use coefficients of balanced reaction equation to convert **moles** of known material **to moles** of desired material.
- Conversion Factor # 3: Use molar mass to convert **moles** of desired material **to mass** of desired material.

Set up as Dimensional Analysis Problem and solve with one equation (not three)!

How many grams of oxygen are produced when 4.0 g butane (C_4H_{10}) combusts?

Figure 7.1.2 solves this problem with one equation, where each step is a different conversion factor (equivalence statement). Each of these steps is detailied in the bullet list below

- Conversion Factor #1: **Mass** $\rightarrow$ **Mole**. Divide by molar mass of butane. Note, this has significant digits, and the units are included in the numerator and denominator as they are different, but the identity is the same and so not repeated in the denominator (it is implicitly understood that it is the same chemical entity, which in this case is butane)
- Conversion Factor #2: **Mole**$\rightarrow$ **Mole**. Use coefficients of balanced reaction equation to convert moles of known material to moles of desired material. These are exact numbers that come from the balanced equation and so do not have significant Figures, but since they are relating two different entities, the identity of both the numerator and denominator are explicitly stated (oxygen in the numerator, butane in the denominator).
- Conversion Factor # 3: **Mole**$\rightarrow$ **Mass**. Multiply by molar mass to convert moles oxygen to grams oxygen. Note, this has significant digits, and the units are included in the numerator and denominator as they are different, but the identity is the same and so not repeated in the denominator (it is implicitly understood that it is the same chemical entity, which in this case is oxygen butane)

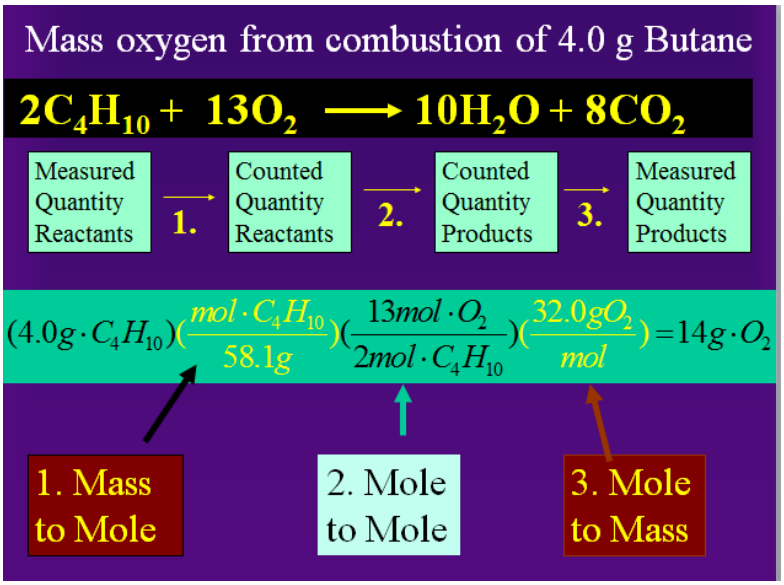


Figure 7.1.2: Solving a stoichiometric problem using all three conversions in one equation.

Cancel units in the above equation to be sure your answer has the desired units and only round off your final answer to the correct number of significant digits, do not round off during intermediate steps.

How many significant digits should your molar masses be? This can vary from instructor to instructor, but the rule of thumb is to use one more significant digit than the measured values of your problem statement. Since this problem had 4.0 g of butane, which has 2 significant digits, we used molar masses of 3 significant digits, but rounded the final answer to two. If at all possible, never use a molar mass with fewer significant digits than the measured values you are working with, as stated in the problem.

Do you Always start with the Mass-to-Mole step? *NO!* If the question was what mass of oxygen is produced if you combust 4 moles of butane, you do not need to do the first step. Note, both propane and carbon dioxide are gases at ambient conditions and so you can not measure their mass directly. Instead, you need to measure the pressure, temperature and volume of the gas to determine the number of moles. We will learn how to do this when we study gasses, but for now we will use masses in our calculations.

Worked Problems

Consider the formation of Analine ($C_6H_5NH_2$) from from nitrobenzene ($C_6H_5NO_2$), iron and water.

$$4C_6H_5NO_2 + 9Fe + 4H_2O \rightarrow 4C_6H_5NH_2 + 3Fe_3O_4 \tag{7.1.4}$$

Lets look at some of the questions you could answer. (AND BE CAREFUL TO READ THE QUESTIONS!)

1. How many moles of iron would you need to produce 100 mole of analine?
2. How many grams of iron would you need to produce 100 mole of analine?
3. How many grams of iron would you need to produce 100 grams of analine?
4. How many grams of analine could 100 g nitrobenzene produce?
5. How many moles of Fe_3O_4 could 100 g of nitrobenzene produce?
6. How many moles of Fe_3O_4 would be produced if 100 g of analine was produced?
7. ...and many more permutations....SO READ THE QUESTION CAREFULLY!

Table 7.1.1; molar masses of chemicals in Equation 4.1.2

$C_6H_5NO_2$	Fe	H_2O	$C_6H_5NH_2$	Fe_3O_4
123.105g/mol	55.845g/mol	18.016g/mol	93.121g/mol	231.535 g/mol

Mol-to-Mol

Question 1: How many moles of iron is needed to produce 100.0 mol of Analine?

Implicit in this question is that nitrobenzene and water are in excess. **Note: a good technique is to write down the equation and under each species identify what is given, and what is being asked**

$$4C_6H_5NO_2 + \underbrace{9Fe}_{moles=?} + 4H_2O \rightarrow \underbrace{4C_6H_5NH_2}_{100\ \text{moles}} + 3Fe_3O_4 \tag{7.1.5}$$

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Video 7.1.1: 1'25" YouTube solution to question 1 (https://youtu.be/Rt5S-wQ0FD8)

$$100\ \cancel{mol\ C_6H_5NH_2}\left(\frac{9\ mole\ Fe}{4\ \cancel{mole\ C_6H_5NH_2}}\right) = \tag{7.1.6}$$

Mass to Mass

Question 2: How many grams of analine could 100 g nitrobenzene produce?

We are now solving a problem where we can actually measure the quantity of reactants and products involved. We start by writing an equation, and under each species write the given values, while identifying what we are solving for.

$$\underset{\substack{mass=100.0g\\ fw=\frac{123.105g}{mol}}}{4C_6H_5NO_2} + 9Fe + 4H_2O \rightarrow \underset{\substack{mass=?g\\ fw=\frac{93.121g}{mol}}}{4C_6H_5NH_2} + 3Fe_3O_4 \tag{7.1.7}$$

We were given 100 g of nitrobenzene and wanted the mass of analine that could be produced, so the first thing I did was write that down under each species. I also needed the molar mass of nitrobenzene and analine, and so wrote that under those. Now on the equation I have everything I need to solve (Video 7.1.2

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Video 7.1.2 3'03" Youtube solving question 2 (https://youtu.be/WQLgWy8O0u0)

$$100\ \cancel{gC_6H_5NO_2}\frac{4\ \cancel{mole\ C_6H_5NO_2}}{123.105\ \cancel{g}}\left(\frac{4\ \cancel{mol\ C_6H_5NH_2}}{4\ \cancel{mol\ C_6H_5NO_2}}\right)\frac{93.121g\ C_6H_5NH_2}{\cancel{mol}} = \tag{7.1.8}$$

Mass to Mole

Question 3: This time we want to know how many moles Fe_3O_4 would be formed from 100 g nitrobenzene. We start by writing down what we are given and what we are solving for (100g nitrobenzene given and moles of Fe_3O_4 sought for) and below that we right any additional information we need to solve this problem. That is, we need the molar mass of nitrobenzene in order to calculate the moles. We do not need the molar mass of Fe_3O_4 because the question is looking for how many moles, not how many grams.

$$\underbrace{4C_6H_5NO_2}_{\frac{mass=100g}{fw=\frac{123.105g}{mol}}} + 9Fe + 4H_2O \rightarrow 4C_6H_5NH_2 + \underbrace{3Fe_3O_4}_{moles=?} \tag{7.1.9}$$

YouTube Video

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Video 7.1.3: 2'23" YouTube solving question 3 (https://youtu.be/kaE2FaFk260).

$$100\ g\ \cancel{C_6H_5NO_2}\ \frac{\cancel{mole\ C_6H_5NO_2}}{123.105\ \cancel{g}}\left(\frac{3\ mol\ Fe_3O_4}{4\ \cancel{mol\ C_6H_5NO_2}}\right) = ? \tag{7.1.10}$$

Test Yourself

Homework: Section 7.1

Graded Assignment: Section 7.1

Exercise 7.1.1

What is true regarding the following statements made for the following balanced chemical reaction?

4 Al(s) + 3 O_2(g) --> 2 Al_2O_3(s)

A) for every 8 moles of aluminum and 6 moles of oxygen consumed, 4 moles of aluminum oxide are produced.

B) for every 4 molecules of aluminum and 3 molecules of oxygen consumed, 2 molecules of aluminum oxide are produced.

C) for every 4 grams of aluminum and 3 grams of oxygen consumed, 2 grams of aluminum oxide are produced.

D) A and B only

E) A, B, and C

Answer available at the end of this section.

Contributors and Attributions

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both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Anonymous
- Ronia Kattoum

Answers

D) A and B only. Stoichiometric ratios in a balanced chemical equations represent a particles/particle ratio that can be scaled up to a mole/mole ratio. So, instead of looking at one molecule at a time, which is not very feasible, we can count by the mole (think by the dozen) of particles at time. Coefficients never represents a mass/mass to ratio because each species in a chemical reaction has a different molar mass. Note: A is still true. We are just doubling the amount of moles, but the ratio is still the same.

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7.2: Theoretical Yield, Limiting and Excess Reagents

Learning Objectives

- Understanding Stoichiometric Proportions
- Understanding Limiting and Excess Reagents
  - Predict quantities of products produced or reactants consumed based on complete consumption of limiting reagent (on both mole and mass basis)
  - Predict quantities of excess reagents left over after complete consumption of limiting reagents.

Stoichiometric Proportions and Theoretical Yield

A balanced chemical equation describe the ratios at which products and reactants are respectively produced and consumed. That said, the coefficients of the balanced equation have nothing to do with the actual quantity of reactants you start with, as you can mix any amount you choose, but clearly the maximum yield **(theoretical yield)** must be limited by the reactant that gets consumed up first, the **limiting reagent**. Any reagents remaining after the complete consumption of the limiting reagent are know as **excess reagents.**

YouTube Video

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Video 7.2.1 : 4/34" YouTube uploaded by "Professor Dave Explains", the first 2:47 goes over Limiting and excess reagents, and then it goes into Percent Yield (section 4.2.3)

Here is some common terminology used to describe reactions based on the concentrations of reactions.

- **Stoichiometric Proportions:** Reactants are mixed in the ratios defined by their stoichiometric coefficients. If a reaction proceeds to completion, everything is consumed.
- **NonStoichiometric Proportions**: Reactants are mixed in ratios that are different than the stoichiometric coefficients. One species runs out first (Limiting Reagent), while another is not completely consumed (Excess Reagent).
  - **Excess Reagent:** The quantity (mole or mass) left over after the complete consumption of the limiting reagent

    Quantity Excess = Initial Quantity - Consumed Quantity.

- **Theoretical Yield**: the maximum possible yield based on the complete consumption of the limiting reagent

Thought Question 7.2.1

Are the limiting reagents always completely consumed to produce products?

Answer available at the end of this section.

Limiting Reagent Problem Strategies:

1. **Identify moles of all reactants present.**
   If given mass, divide by formula weight to convert to moles (this is the mass to mole step from the section 4.1,3.
2. **Divide moles of each reactant by it's stoichiometric coefficient.**
   This is the denominator of the mole-to-mole step in section 4.1.3.
3. **Smallest number indicates limiting reagent.**
4. **Multiply by stoichiometric coefficient of species you are solving for, and answer the question .**
   This is the numerator of the mole-to-mole step in section 4.1.3. If you are after moles, you are finished, if you are after mass, you need to use the molar mass of product to convert moles product to grams mass product, which is the mass-to-mole step in section 4.1.

Limiting Reagent Problem

Silver tarnishes in the presence of hydrogen sulfide and oxygen due to the following reaction

$4Ag + 2H_2S + O2 \text{ ----> } 2Ag_2S + 2H_2O$

What is the Limiting Reagent and Theoretical Yield of Ag_2S if 2.4 g Ag, 0.48 g H_2S and 0.16g O_2 react?

Note in the video how we first wrote the balanced equation, and then under each species wrote down what we were given. In this problem there are 3 reagents, and this technique allows us to quickly identify the

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Video 7.2.2: 6'07" YouTube calculating the theoretical yield of Ag_2S based on the complete consumption of the limiting reagent (https://youtu.be/26ieM_k0NOU) [Digital Corrigendum: @1'50" in the video the mass silver divided by its molar mass and stoiochiometric coefficient was written down to be 0.0556 but was 0.00556. at 2'22" the value of oxygen was determined to be 0.005, and so that was smaller and was the limiting reagent. Note: at 2:57 this problem is redone the long way, and this allows you to see how the shortcut works.

Excess Reagent

To calculate the excess reagent you determine how much is left over after the complete consumption of the limiting reagent

$$\text{Mass}_{\text{excess reagent}} = \text{Mass}_{\text{initial}} - \text{Mass}_{\text{consumed by complete consumption of limiting reagent}}$$

Now building on the last problem, calculate the mass of excess reagents for the reaction of 2.4 g Ag, 0.48 g H_2S and 0.16g O_2 to form $Ag_2S + H_2O$.

YouTube Video

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Video 7.2.3: 4'37" YouTube calculating the excess reagent concentrations when 2.4 g Ag, 0.48 g H_2S and 0.16g O_2 react.

Test Yourself

Homework: Section 7.2

Graded Assignment: Section 7.2

Exercise 7.2.1

Consider the oxidation of glucose through respiration:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + Energy \tag{7.2.1}$$

What mass of carbon dioxide forms when 25.00 g of glucose reacts with 40.0 g of oxygen?

Answer available at the end of this section.

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- Ronia Kattoum (UALR)
- anonymous

Answers

Thought Question 7.2.1

No, only if the reaction goes to completion. There can be many different reasons why the limiting reagent is not completely consumed, these can include:

- Products also react to form reactants causing an equilibrium of reactants of products to coexist, this will be covered next semester (see Chapter 15)
- Reactions may not be over (some reactions occur very slowly).

In some cases the consumption of the limiting reagent may not lead to product production.

- Intermediates form and are not completely consumed
- Secondary (other) reactions occur

Step 1: Calculate moles of each reactant:

$$25.00\ \text{g} \times \frac{1\ \text{mol}}{180.06\ \text{g}} = 0.1388\ \text{mol}\ C_6H_{12}O_6$$

$$40.0\ \text{g} \times \frac{1\ \text{mol}}{32\ \text{g}} = 1.25\ \text{mol}\ O_2$$

Step 2: Divide each by its stoichiometric coefficient, smallest value is limiting reagent.

$$1.25 mol O_2 (\frac{1}{6 mol}) = 0.208 \tag{7.2.2}$$

$$0.1388 mol C_6H_{12}O_6 (\frac{1}{1 mol}) = 0.1388$$

So glucose is the limiting reagent.

Step 3: calculate the mass carbon dioxide based on the complete consumption of the limiting reagent.

$$0.1388mol\ C_6H_{12}O_6(\frac{6mol\ CO_2}{1mol\ C_6H_{12}O_6})\left(\frac{44.011g\ CO_2}{mol}\right) = 36.66g\ CO_2 \tag{7.2.3}$$

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7.3: Percent Yield

Learning Objectives

- Differentiate between actual yield, theoretical yield, and percent yield
- Calculate percent yield

Percent Yield

Percent Yield is defined as the actual yield divided by the theoretical yield times 100.

$$\text{Percent Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right) \times 100\% \tag{7.3.1}$$

There are many reasons why the actual yield of a chemical reaction may be less than the theoretical yield, and these will be taken up during later Chapters of the course. Here are some reasons, most of which deal with topics that will be covered in the second semester of this course.

1. Equilibria between products and reactants, where the limiting reagent is not completely consumed (Chapter 15)

$$Ca_3PO_4(s) \leftrightharpoons 3Ca^{+2} + 2PO_4^{-3} \tag{7.3.2}$$

2. Kinetics, simply speaking, the reaction may be very slow and not over (Chapter 14).
3. Formation of Intermediates. A chemical equation does not represent the mechanism but the overall balance of mass. The mechanism tells how the reaction actually proceeds (section 14.6), and there are often intermediate chemical species that are not completely consumed. This is exemplified in the following reaction.

$$H_2(g) + 2ICl(g) \rightarrow I_2(g) + 2HClg \tag{7.3.3}$$

It is highly unlikely that this reaction takes place in one step where three molecules simultaneously collide. A more plausible way for the reaction to occur is in two steps as below, where HI is an intermediate.

$$H_2 + ICl \rightarrow HCl + HI \tag{7.3.4}$$
$$HI + ICl \rightarrow HCl + I_2 \tag{7.3.5}$$

In the first step HCl and the intermediate HI are formed from the bimolecular collision of H_2 and ICl. In the second step, the HI (intermediate) collides with another ICl to form HCl and I_2. If you add the two steps the intermediate cancels out and you get the balanced equation.

$$H_2(g) + 2ICl(g) \rightarrow I_2(g) + 2HClg \tag{7.3.6}$$

Percent Yield Problem

Phosphoric Acid can be synthesized from Phosphorous, Oxygen and water according to the following equations

$$4P + 5O_2 + 6H_2O \rightarrow 4H_3PO_4 \tag{7.3.7}$$

What is the percent yield if 50.3 g of phosphoric acid is created from 20.0 g phosphorous, 15.0 g of water and 30.0 g of oxygen?

First we set up the problem and it is suggested that we write down the given quantities under each species of the balanced equation (as that makes it easier to see the given information)

$$\underset{\substack{m=20.0g\\ fw=30.98g/mol}}{4P} + \underset{\substack{m=30.0g\\ fw=32.00g/mol}}{5O_2} + \underset{\substack{m=15.0g\\ fw=18.01g/mol}}{6H_2O} \rightarrow \underset{\substack{ActualYield=50.3g\\ fw=97.995g/mol\\ TheoreticalYield=?}}{4H_3PO_4} \tag{7.3.8}$$

After calculating the theoretical yield (based on the complete consumption of the limiting reagent) we calculate the percent yield. This problem is solved in video 7.3.1.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

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Video 7.3.1: 4'40" YouTube solving a percent yield problem (https://youtu.be/BCOdyr1QKj0).

Exercise 7.3.1

A chemist decomposes 1.006 g of $NaHCO_3$ and obtains 0.434 g of Na_2CO_3. What are the theoretical yield and the actual yield? What is the percent yield?

$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(\ell) + CO_2(g)$

Answer available at the end of this section.

Test Yourself

Contributors and Attributions

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Answers

theoretical yield = 0.6345 g; actual yield = 0.434 g
percent yield = 68.4%

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8: SOLUTION CHEMISTRY



- Topic hierarchy

8.1: Measuring Concentrations of Solutions
8.2: Acid and Base Concentrations-pH scale
8.3: Stoichiometry of Aqueous Reactions

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8.1: Measuring Concentrations of Solutions

Learning Objectives

- Define Molarity
- Differentiate between solute, solvent, and solution
- Calculate the molar concentration for solutes
- Outline the steps to make a solution of a desired concentration from a solid or aqueous solute
- Calculate the concentration of ions in a soluble ionic compound
- Perform stoichiometric calculations involving aqueous solutes
- Calculate the concentration of unknown solutes

Introduction

Up to this point we have used stoichiometry to "count" atoms, molecules and ions by measuring mass of pure substances and using molar masses to calculate the number of chemical entities (moles). Many of the reactions we have studied involve solutions, where we are interested in a solute that is dissolved in a solvent, and we can not measure the mass of the solute independent of the solvent. We also need to realize that many chemical reactions require the reactants to be mobile, where they can bump into each other, and this can occur when they are dissolved in a solvent. So it is very important that we can count chemical entities when the entity of interest is a solute dissolved in a solvent, and in which case we measure the mass of volume of the solution as a whole (solvent plus all solutes) of which volume is typically the easiest to measure.

Concentration of a Solute

There are two basic ways of reporting the concentration of a solute in a solvent, by reporting the mass of solute in a given volume, or the number of moles of solute in a given volume. These are effectively conversion factors that define the equivalent mass or moles of a solute to the volume of the solution.

- **Mass Concentration**: has typical units of g/L

$$\frac{mass\ solute}{vol\ solution} = \frac{m(g))}{V(L))} \tag{8.1.1}$$

- **Mole Concentration (Molarity)**: has units of mol/L or M . You can consider M as a shorthand notation for mol/L.

$$\frac{moles\ solute}{vol\ solution} = \frac{n(moles))}{V(L))} = M\ or\ (mol/L) \tag{8.1.2}$$

So a 3.0M solution of sucrose has 3.0 mole of sucrose per liter

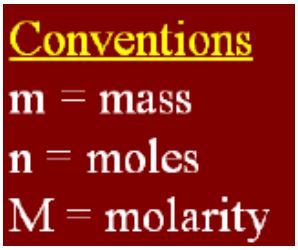


Figure 8.1.1: In order to avoid confusion we will use a small "m" to denote mass, and a large ""m" to denote Molarity.

How do we convert between mass and mole based concentration?

Simply by multiplying or dividing by the molar mass (g/mol). If you know the g/L, you simply divide by the molar mass, and if you know the moles/L and want the mass/L, you multiply by the molar mass (just use the molar mass as an equivalence statement).

How do we "count" solute molecules and ions?

By measuring the volume and knowing the molarity. This is analagous to measuring the mass and knowing the molar mass for a pure substance.

$$n(moles) = M(\frac{mol}{L})(V(L)) \qquad (8.1.3)$$
$$n = MV$$

This can be contrasted to how we "count" chemical entities that are pure solids.

$$n(moles) = \frac{m(g)}{molar\ mass(\frac{g}{mol})} \qquad (8.1.4)$$
$$n = \frac{m}{fw}$$

Making a Solution with a Solid Solute.

Step 1: Calculate Mass of solute needed for desired volume
Step 2: Quantitatively Transfer Mass to Volumetric Flask
Step 3: Dilute to volume with solvent, ensuring that all of the solid has dissolved.

Making 500.0 mL of 0.500 M Copper(II)Sulfate

Step 1: Calculate Mass $CuSO_4$(s) needed.

$$0.5000L\left(\frac{0.500molCuSO_4}{L}\right)\left(\frac{159.6gCuSO_4}{mol}\right) = 39.9gCuSO_4 \qquad (8.1.5)$$

Step 2: Weight 39.9g $CuSO_4$(s) and quantitatively transfer to 500 mL graduated cylinder (which is calibrated to 500.0 mL), being sure all the salt is transferred.

Step 3: Fill half way and mix, then dilute to volume. Make sure all the solute is dissolved, and recheck that solution level is at bottom of meniscus.

Exercise 8.1.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

Why can't you just add 500.0mL of water to 39.9 g $CuSO_4$ to make 500.0 mL of 0.500M solution?
First, this is not taking into account the volume of the solute, and second, there is usually a change in volume when you dissolve a solute into a solvent, and in the case of ionic compounds this is a contraction.

Exercise 8.1.2

Using concentration as a conversion factor, perform the following calculation.

a. How many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?
b. What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?
c. What volume of 1.50 M HCl solution contains 10.0 g of hydrogen chloride?

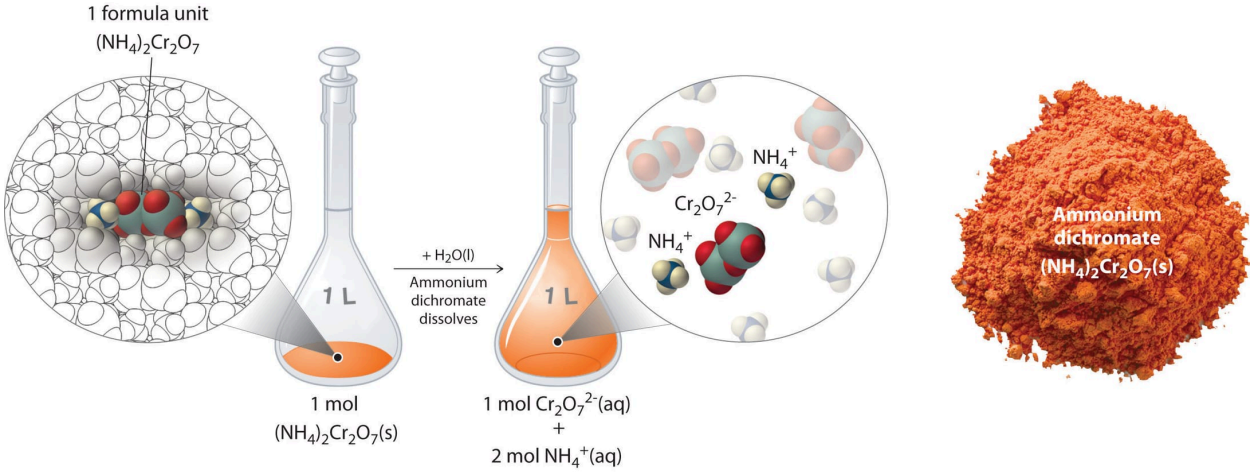
Answer a available at the end of this section.

Ion Concentrations in Solution

When an ionic compound dissolves it breaks up into its ions.

$(NH_4)_2Cr_2O_7(aq) \rightarrow 2NH_4^+(aq) + Cr_2O_7(aq)$

Dissolving 252.07 grams of ammonium dichromate (fw = 252.07 g/mol) results in a solution that is 1M ammonium dichromate, but when an ionic salt dissolves, it breaks up into ions, and so what you really have is a solution that is 2 M ammonium ion (NH_4^+) and 1 M in dichromate ion ($Cr_2O_7^{+2}$)



Dissolution of 1 mol of an Ionic Compound. In this case, dissolving 1 mol of $(NH_4)_2Cr_2O_7$ produces a solution that contains 1 mol of $Cr_2O_7^{2-}$ ions and 2 mol of NH_4^+ ions. (Water molecules are omitted from a molecular view of the solution for clarity.)

Example 8.1.1 Solution Concentration of Iron(II)chloride

Water is added to 2.16 g of the ionic compound ferrous chloride to make a solution with a total volume of 100.0 mL. Express the concentration of the salt solution, and that of its ions.

1. What is the salt concentration?
2. What are the ion concentrations?

Solution

Video 8.1.1 goes over this calculation.

Ans: 0.170M $FeCl_2(aq)$

0.170M $Fe^{+2}(aq)$

0.340M $Cl^-(aq)$

Yes, there are three answers! (that of the salt, and of the ions)

YouTube Video

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Video 8.1.1: 3'36" YouTube calculating the concentration of a Ferrous Chloride in example 8.1.1.

Exercise 8.1.3

You have a 1.50 M solution of Na_2CO_3. What is the concentration of:

A) sodium ions?

B) carbonate ions?

C) total ions?

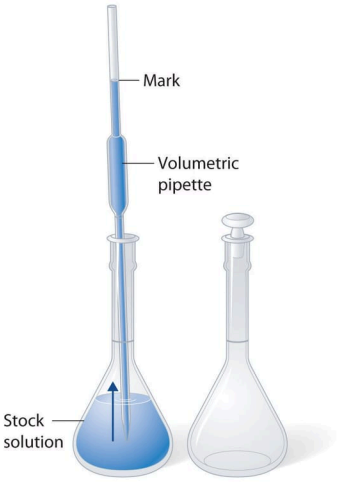
Answer available at the end of this section.

Preparing Solutions with Stock Concentrated Solutions (Dilution).

Many stockroom reagents come as concentrated solutions that can easily be diluted to a desired concentration by adding solvent. Adding a solvent does not change the moles of solute (n), so

$$n_{initial} = n_{final}$$
$$n_i = n_f$$
$$M_iV_i = M_fV_f$$

Step 1: Calculate initial volume of stock
Step 2: Transfer to volumetric flask with a volumetric pipette
Step 3: Dilute to Volume



(a) A volume (V_s) containing the desired moles of solute (M_s) is measured from a stock solution of known concentration.



(b) The measured volume of stock solution is transferred to a second volumetric flask.



(c) The measured volume in the second flask is then diluted with solvent up to the volumetric mark [$(V_s)(M_s) = (V_d)(M_d)$].

Example 8.1.2

How do you make 500.0 ml of 0.70M HCl solution from stock 11.6 M HCl?

Solution

Step 1: Calculate initial volume of 11.6 M hydrochloric acid needed.

$$M_i V_i = M_F V_F \tag{8.1.6}$$

$$V_i = V_f \left(\frac{M_f}{M_i}\right) = 500.0mL \left(\frac{0.70M}{11.6M}\right) = 30.2mL \tag{8.1.7}$$

Step 2: Quantitatively transfer this volume to a 500 mL volumetric flask (this has 4 significant Figures, that is, it is calibrated to 500.0 mL)

Step 3: Dilute with water to mark

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Video 8.1.2: Dilution problem on creating 500.0 ml of a 0.700M HCl solutioin

Exercise 8.1.3

What is the concentration of an HCl solution if 598 mL of 0.778 *M* HCl is diluted to 1.00 L?

Answer available at the end of this section.

Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- November Palmer & Ronia Kattoum (UALR)
- anonymous

Answers

1.57 M $C_6H_{12}O_6$

1.74 L

$$0.0773mol\ CH_2O\left(\frac{L\ solution}{0.0444mol\ CH_2O}\right) = 1.7409L \tag{8.1.8}$$

6.12 g

$$1.08L\ solution\left(\frac{0.0578mol\ H_2SO_4}{L\ solution}\right)\left(\frac{98.08g\ H_2SO_4}{mol}\right) = 6.122546g\ H_2SO_4 \tag{8.1.9}$$

183 mL or 0.183L

$$10.0g\ HCl\left(\frac{mol\ HCl}{36.46g}\right)\left(\frac{L\ solution}{1.5mol\ HCl}\right) = 0.182849L \tag{8.1.10}$$

A) $[Na^+]$ = 3.00 M

B) $[CO_3^{-2}]$= 1.50 M

C) $[Na^+]$ + $[CO_3^{-2}]$ = 4.50 M ; or you can choose think for every 1 mole of Na_2CO_3, there are 3 moles of ions (identify of the ions is irrelevant).

$$M_iV_i = M_FV_F$$

$$M_f = M_i\left(\frac{V_i}{V_f}\right) = 0.778M\left(\frac{0.0598L}{1.00L}\right) = 0.465M$$

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8.2: Acid and Base Concentrations-pH scale

The pH Scale

What is pH?

The pH scale measures the concentration of hydronium ions in water. In section 3.5 we looked at the autoionization of water, which is when two water molecules collide and one gives a proton to another, resulting in the formation of hydronium and hydroxide ions.

$$H_2O(l) + H_2O(l) \text{ --> } H_3O^+(aq) + OH^-(aq)$$

For neutral water the hydronium and hydroxide concentrations are equal, and this is a very small number, 10^{-7}M. As these are very small numbers we use a log scale to represent pH, which is a convenient way of expressing the hydronium ion concentration of a solution. It is fairly common knowledge that neutral water has a pH of 7, acids have a pH <7 and bases have a pH>7, but few people understand this in terms of the actual hydronium ion concentration. Our objective is to develop an understanding of logarithmic scales by developing a pH scale. First we will introduce the **convention of using brackets to describe concentration**, so $[H_3O^+]$ means the concentration of hydronium in moles per liter (M).

$$\mathbf{pH = -\log[H_3O^+]}$$

$$\mathbf{[H_3O^+] = 10^{-pH} = 1/10^{pH}}$$

So a $[H_3O^+]$ concentration of $1.0x10^{-7}$ M solution has a pH of 7.00. Note, in pH the significant Figures are given by the number of places to the left of the decimal, as it is not a number, but the number of times something has been divided (or multiplied) by 10. To understand this, consider a solution with a concentration of $1x10^{-12}$, the number has 1 significant Figure and the pH is 12. Since 12 has two digits, we can't use the digits to the left of the decimal, and so we write it as 12.0, indicating it has 1 significant digit.

To get a feel for this we can use the following applet and the method of successive dilutions to develop a pH scale.

Developing pH Scale via Method of Successive Dilutions

Method of Successive Dilutions:

We have created a demonstration video, which describes the features and interface of the Virtual Lab.

In the virtual lab below:

Part 1: Setting up the lab for the method of successive dilutions.

- Open the stockroom and go to the strong acids
  - single click on 1 M HCl,
  - single click on the 3 L distilled water
- Click on "glassware" and move into the workbench a 100 mL volumetric flask
- Drag the water over the 100 mL volumetric flask and transfer 90 mL of water into the flask.
- right click on the volumetric flask and hit "duplicate". repeat this 7 times, so you have 8 100 mL flasks with each containing 90 mL of water.
- Click on 1 M HCl, and read the pH, which will be 0, because $1 = 10^{-pH} = 10^0 = 1$.

Part 2: Successive Dilutions.

- Click on the stock 1 M HCl and record the pH
- Transfer 10 ml of 1 M HCl to a volumetric flask with 90 mL water, so the total volume is 100 mL, and record the pH of the diluted solution, which represents 10 mL osf the stock dilute to 100 mL,
  - From the dilution equation, $M_iV_i = M_iV_i$ and

$$M_f = M_i \left(\frac{V_i}{V_f} \right) \tag{8.2.1}$$

    Note, the ratio of the volumes is unit-less and can be defined as the dilution factor, which in this case is 10:100 or 1 to 10, so the diluted solution is 1/10th the original.
  - If you successively dilute each solution 10 fold by pouring 10 mL into 90 mL of water, the pH of each solution is n, where n is the number of times you successively diluted the stock 1M solution.
  - That is, each solution is $1/10^n$ the concentration of the original solution, where n is the number of times the original solution was successively diluted, and so the pH is the number of 1/10 dilutions.

Why after eight 10-fold dilutions does the pH not become 8, but stays at 7?

Neutral water is the result of one water molecule transferring a proton to another water molecule, and this results in a hydronium concentration of 10^{-7}. When the acid is diluted to a concentration of less than 10^{-7}, it is so dilute that it protonates fewer water molecules than the water does.

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8.3: Stoichiometry of Aqueous Reactions

Introduction

At the beginning of this Chapter we covered stoichiometry of reactions where everything was a solid, and we used the mass and formula weight to calculate the moles of reactants or products consumed or produced. For a reaction where two chemicals interact with each other they must come into contact with each other if the reaction is going to proceed. One way of enabling this mobility is to dissolve the reactants into a solvent creating a solution. Unless you have a pure substance, you can not use mass to measure the number of particles, as you are weighting both the solvent and solute(s). So if you have a solution, you need to know the concentration, and the most common way of representing chemical concentrations is the molarity.

Solid phase:

$$n(moles) = \frac{m(g)}{fw(\frac{g}{mol})} \tag{8.3.1}$$

$$n = \frac{m}{fw} \tag{8.3.2}$$

Aqueous phase (solute concentration):

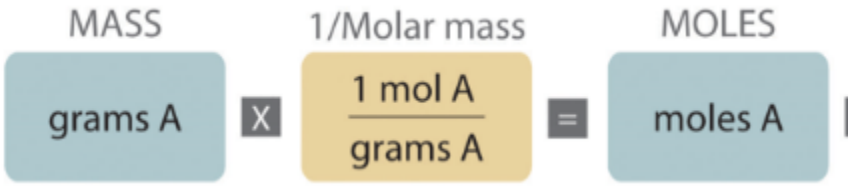
$$n(moles) = M\left(\frac{mol}{L}\right)V(L) \tag{8.3.3}$$

$$n = MV \tag{8.3.4}$$

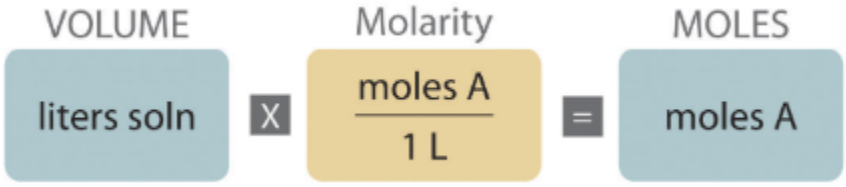
where,
n = moles
m = mass
V = Volume
fw = formula weight (molar mass)

So now we can calculate the moles of a pure substance if we weigh it's mass and know it's molar mass, or if we measure its volume and know it's concentration (Molarity).

If given a solid, you need to know its molar mass



If given a solute, you need to know its concentration (molarity)



Solution Stoichiometry

In sections 4.1 and 4.2 we calculated theoretical yield, limiting reagent and excess reagents for reactions where all chemical

species were described as solids where we measured the mass. Now that we have a way to describe solute concentrations (Molarity) we can extend these to include solutions where we measure the volume. Lets look at the reaction of adding a solid soluble salt to an aqueous solution

Dissolving 10.0 g of Copper(II)sulfate into 500.0 mL of 0.100 M sodium phosphate

Exercise 8.3.1

What are the formulas of the reactants when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate? (see video 8.3.1)

Answer available at the end of this section.

Exercise 8.3.2

What are the formulas and phases of the products when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate?(see video 8.3.1)

Answer available at the end of this section.

Exercise 8.3.3

Write the balanced equation when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate?(see video 8.3.1)

Answer available at the end of this section.

NOTE: We added copper(II)sulfate as a solid, but it dissolved (see solubility rules, section 3.4.2.1), and so it should be treated as aqueous when writing the equation

YouTube Video

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Video: 8.3.1: 3'54" YouTube identifying reactants, products, phases and balancing the equation for the dissolution of 10 g of copper(II) sulfate into 500.0 mL of 0.100 M sodium phosphate (https://youtu.be/qdtiFATDMdU).

Calculating the Theoretical Yield for the mass of precipitate formed

Example 8.3.4

Calculate the theoretical yield for the mass of precipitate formed when 10.0 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate.

Solution

See video 8.3.2 for the solution of this problem, and video 8.3.3 for a deeper explanation of the fast technique used in video 2. Note how we set up the equation by writing out the given data under each species, along with the moles that we need to calculate with that data. We calculate for the moles of each reactant (n), divide by its storichiometric coefficient, and the smallest number is the limiting reagent. We then proceed and calculate the mass of copper(II)phosphate formed.

$$\underset{\substack{n=\\ m=10.0g\\ fw=159.61g/mol}}{3CuSO_4(s)} + \underset{\substack{n=\\ M=0.100mol/l\\ V=500.0ml}}{2Na_3PO_4(aq)} \rightarrow \underset{\substack{m=?\\ fw=380.58g/mol}}{Cu_3(PO_4)_2(s)} + 3Na_2SO_4(aq) \tag{8.3.5}$$

at 2'25" into video 8.3.2 we calculated the limiting reagent, which was copper(II)Sulfate and thus based the yield on its consumption. The overall equation is

$$10.0gCuSO_4(s)\left(\frac{molCuSO_4}{159.69g}\right)\left(\frac{1molCu_3(PO_4)_2}{3molCuSO_4}\right)\left(\frac{380.58gCu_3(PO_4)_2(s)}{mol}\right) = 7.95gCu_3(PO_4)_2(s) \tag{8.3.6}$$

Note, the above equation is not used in the fast technique, but is used in video 8.3.3 when we explain the fast technique

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Video: 8.3.2: 3'30" YouTube video calculating the mass precipitate formed by complete consumption of the limiting reagent for the dissolution of 10 g of copper(II) sulfate into 500.0 mL of 0.100 M sodium phosphate (https://youtu.be/Vh3Je39bgQc)

YouTube Video

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Video: 8.3.3: The fast technique in video 8.3.2 explained in more detail (https://youtu.be/5Qlo1bi1uzE).

Calculating excess reagents

Since the copper sulfate is the limiting reagent we can calculate the amount of excess sodium phosphate

Example 8.3.2: Excess Reagent

Calculate the concentration (molarity) of excess phosphate when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate (note: the real reaction is expressed by the net ionic equation:

$$3Cu^{+2}(aq) + 2PO_4^{-3}(aq) \rightarrow Cu_3(PO_4)_2(s) \tag{8.3.7}$$

Solution

see Video 8.3.4

First calculate moles excess phosphate

$$\text{Initial moles} - \text{amount consumed by reaction with limiting reagent} \tag{8.3.8}$$

$$0.050 molPO_4^{-3} - 0.06265 molCu^{+2}\left(\frac{2\ mol\ PO_4^{-3}}{3\ mol\ Cu^{+2}}\right) = 0.0082 molPO_4^{-3}$$

then divide into total volume

$\frac{0.0082 molPO_4^{-3}}{0.500l} = 0.016 MPO_4^{-3}(aq)$

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Video 8.3.4: Determining the excess phosphate when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate https://youtu.be/xsp1_BADnqA).

Calculating Concentration of Spectator Ions

For the reaction that occurs when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate we have calculated the theoretical yield

Example 8.3.3 Spectator Ion Concentrations

Calculate spectator ion concentrations when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate

Solution

See video 8.3.5 for the solution to this problem. There are several ways you can set this up. In calculating the limiting reagent we determined the moles of each reactant and so that would be a logical place to start

$$\underset{\substack{n=\frac{m(g)}{fw(g/mol)}=\frac{10.0g}{159.61g/mol}\\ n=0.06265mol}}{3CuSO_4(s)} + \underset{\substack{n=MV=(0.100mol/l)(0.5000l\\ n=0.0500mol}}{Na_3PO_4(aq)} \rightarrow Cu_3(PO_4))2(s) + 3Na_2SO_4(aq) \tag{8.3.9}$$

$$\frac{molSO_4^{-2}}{V_{total}} = \frac{0.06265 molCuSO_4\left(\frac{1 molSO_4^{-2}}{molCuSO_4}\right)}{0.500l} = 0.125M \tag{8.3.10}$$

$$\frac{molNa^{+}}{V_{total}} = \frac{0.0500 molNa_3PO_4\left(\frac{3 molNa^{+}}{molNa_3PO_4}\right)}{0.500l} = 0.300M \tag{8.3.11}$$

YouTube Video

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Video 8.3.5: 2'44" YouTube calculating the spectator ion concentration when 10 g of copper(II) sulfate is dissolved into 500.0 mL of 0.100 M sodium phosphate.

Test Yourself

Homework: Section 8.3

Graded Assignment: Section 8.3

Exercise 8.3.4

Consider the reaction of sodium phosphate and barium chloride, where 25.00 ml of 0.8000M Na_3PO_4 is mixed with 25.00ml of 0.1000M $BaCl_2$ (and the final volume is 50.00 ml, that is there is no change in volume upon mixing).

a. What is the limiting reagent?
b. What is the identity and mass of precipitate formed?
c. What is the concentration of the sodium spectator ion?
d. What is the concentration of the excess phosphate?

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

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Answers

$CuSO_4(s)$ and $Na_3PO_4(aq)$

$Na_2SO_4(aq)$ and $Cu_3(PO_4)_2(s)$

$3CuSO_4(aq) + 2Na_3PO_4(aq) \rightarrow 3Na_2SO_4(aq)$ and $Cu_3(PO_4)_2(s)$

$BaCl_2$

$0.5016gBa_3(PO_4)_2\ 0.002500molBaCl_2\left(\frac{1molBa_3(PO_4)_3)}{3molBaCl_2}\right)\left(\frac{601.93gBa_3(PO_4)_2)}{mol}\right)=0.5016gBa_3(PO_4)_2$

0.1200M

$$M_fV_f=M_iV_i\ M_F=\frac{M_iV_i}{V_f}\ \frac{0.8000mol/L\ Na_3PO_4\left(\frac{3molNa^+}{molNa_3PO_4}\right)0.02500L\ Na_3PO_4}{0.02500L+0.02500L}=1.200MNa^+$$

$$mol_{excess}PO_4^{-3}=mole_{initial}-mol_{consumed}$$

$$0.8000molNa_3PO_4/L\left(\frac{1molPO_4^{-3}}{molNa_3PO_4}\right)0.02500L-\left(0.100mol/L\ BaCl_2\left(\frac{1molBa^{+2}}{molBaCl_2}\right)0.02500L\left(\frac{2molPO_4^{-3}}{3molBa^{+2}}\right)\right)=0.01833molPO_4^{-3}$$

$$Concentration=\frac{mol_{excess}PO_4^{-3}}{V_T}=\frac{0.01833molPO_4^{-3}}{0.02500L+0.02500L}=0.3667M\ PO_4^{-3}$$

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9: GASES

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- Topic hierarchy

9.1: Gas Properties

9.2: Gas Laws

9.3: Two State Ideal Gas Problems

9.4: Gas Phase Stoichiometry

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9.1: Gas Properties

Introduction

n this Chapter we will take a look into the gas phase. We will start with covering basic gas properties and then learn several "equations of state", which are mathematical equations that relate measurable values like the volume, pressure and temperature of a gas to the quantity of particles (n) in the gas.

Characteristics of Gases

1. **Fluid** - A gas is a fluid and will take the shape of a container it fills. There is zero viscosity (resistance to flow) and it fills the entire container.
2. **Compressible** - The density of a gas changes as its volume changes. Thus, unlike incompressible solids or liquids, density alone can not be used to help identify a gas. That is, liquid water has a density of 1g/ml at normal temperature and pressure, and so that can be used to help identify a substance as water. But gaseous water (water vapor) fills a container, its density depends on the size of the container.
3. **Gases Form Solutions** - Gas forms homogenous mixtures. Note, for liquid phase solutions we often identify a solvent and a solute, where the solute is dissolved in the solvent to form a solution. We usually consider the solvent to be the substance in greatest quantity, and an aqueous solution is one where water is the solvent. This is because the "space" of matter is being occupied by the liquid particles (which is why they are incompressible). In the gas phase, most of the "space" of matter is the "void" (which is why gasses are compressible), and so it is not normal to consider gas phase solutions to have a solute and solvent. Instead we use the concept of mole fraction to identify the relative proportions of different species in a gas mixture.

Measuring Gases: State Variables

State Variables - an **equation of state** relates measurable state variables to the number of gas particles. The **Ideal Gas Law** (**PV=nRT**) is an equation of state that relates the measurable state variables of **Pressure**, **Temperature** and **Volume** to the number of particles. There are other equations of state and we will discuss some at the end of the Chapter, but this course will expect a mastery of the ideal gas equation of state.

PV=nRT
P=Pressure, V=Volume, n=moles, T=Temperature (Kelvin)
R=0.08206(l-atm/mol-K) = 8.314J/(mol-K)

Pressure

Pressure is defined as force per unit area,

$$P = \frac{F}{A} \tag{9.1.1}$$

and the SI unit is the Pascal.

$$Pa = \frac{1N}{m^2} = \frac{1kg}{m \cdot s^2} \tag{9.1.2}$$

Note, the above definition if founded on the SI base units (kg, s and m) and since May of 2019 these have been defined by the 7 defined constants of nature (section 1B.3.1.2), with the relevant units being the Cs hyperfine transition frequency (atomic clock), speed of light in a vacuum and Boltzmann's constant. Of course we will use more practical definitions here, and the important thing to note is any measurable quantity can be traced back to the 7 SI Base units

This pressure of a gas can be understood by visualizing gas particles in constant motion, as exemplified in Figure 9.1.1.

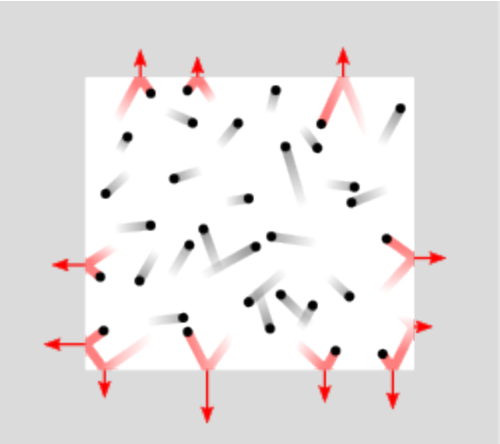


Figure 9.1.1: *Gas particles in a sealed container. As the gas particles move they eventually hit the surface and impart a force on the surface as they change their direction of motion (red particles). The pressure is this force per unit area due to these collisions, which in SI units is the Pa ($1N/m^2$).*

It is more common to see pressure described in atm or torr (mm Hg), which are the units we will predominantly use in this class, and have macroscopic definitions that are more easily measured. Table 10.1.1 shows the common units of pressure, and their conversion factors.

Table 9.1.1: Common units of pressure and conversion factors. UALR students are required to memorize the conversion factors in red

1 atm =	**760 torr**
1 atm =	**760 mm Hg**
1 atm =	14.7 lb/in^2
1 atm =	101,325 Pa

Note, because all values in table 9.1.1 are set to one atmosphere, you can go directly between the other units (14.7 lb/in^2 = $1.01x10^5$ Pa = 760 torr...)

Atmosphere (unit)

One atmosphere of pressure is roughly the pressure felt at sea level by earth's atmosphere, as is shown in Figure 9.1.1. Note, this is not an easily reproducible value because the force felt on the earth changes over both your location and time. This YouTube from the National Naval Aviation Museum does a good job of showing how the air pressure changes as a function of altitude. That is, at the top of Mt Everest (about 5.5 miles high) there is a smaller column of air above you and so the pressure is less than at sea level, while in the Salton Sea CA (-228 ft) the pressure is higher than at sea level.

But also, because the atmosphere of the earth is really a mixture of gasses (mostly nitrogen and oxygen), with many species like water vapor, carbon dioxide, and air pollutants varying over time, so the mass of the column changes over time. So although the concept of an atmosphere of pressure is easy to understand, and it is important because reactions that occur on the surface of the earth occur at around one atmosphere of pressure, its variance over time and location makes it impractical to develop a scale on.

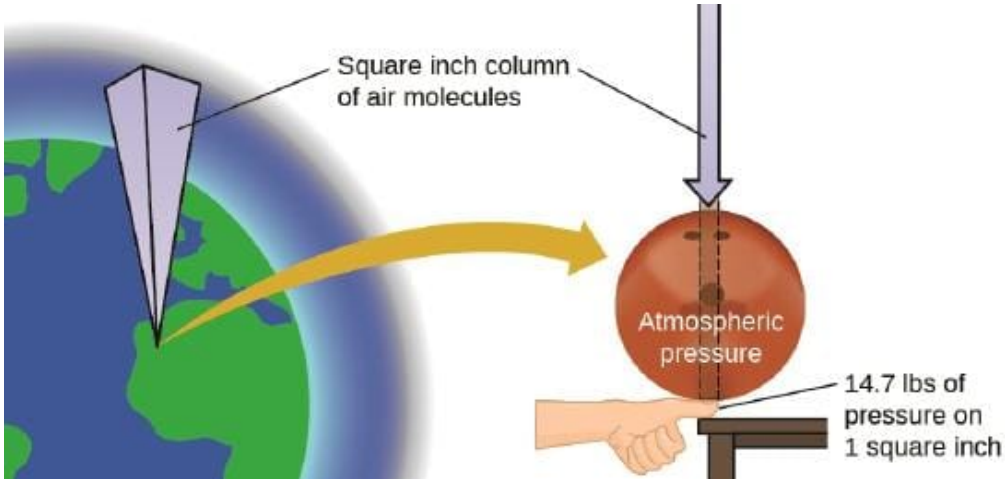


Figure 9.1.2: *This image illustrates the concept of an atmosphere of pressure, which is the pressure felt at sea level due to the column of earth's atmosphere.*

YouTube Video

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View this YouTube Video online

Video 9.1.1: YouTube from State Climate Office of NC demonstrating how the force of atmospheric pressure can collapse an empty can.https://youtu.be/QVayky_b-6U

In the above video, water vapor displaced the air when the water was boiled on the hotplate. It was then inverted and placed in an ice bath such that the opening was under the water surface, this caused the water vapor to condense and a vacuum was formed since there was no way for air to enter, and so the pressure on the inside became less than on the outside, and at some point it became great enough to collapse the can.

Torr (unit)

The torr or mm Hg is a reproducible definition of pressure based on the height of a column of mercury, and can be used to make an instrument to measure atmospheric pressure, the barometer. Using the conversion factor of 1 atm = 760 mm Hg, we could calibrate the barometer in units of atm. This is illustrated in Figure 10.1.3, where the height of the mercury column is dependent on the atmospheric pressure. If the atmospheric pressure is one atm the column is exactly at 760 mm Hg (760 torr). On a humid day the pressure outside goes up and the system is no longer at equilibrium, which forces the mercury up the tube (greater than 760 torr) until they balance out. On an arid day it lowers until they balance out.

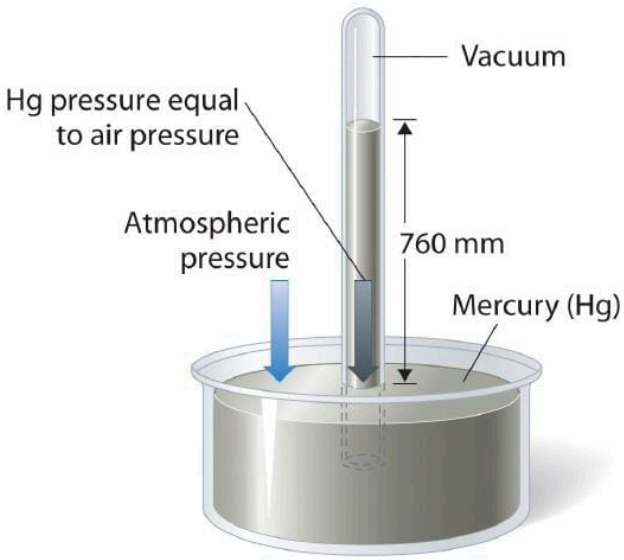


Figure 9.1.3: When at equilibrium the pressure due to the height of the column of mercury must equal the pressure due to the column of air.

Figure 9.1.3 gives a feel for how pressure can be measured. If the atmospheric pressure is higher than that due to the column of mercury it will push the column up until they are equal, if it is less, the column of mercury will go down until they are equal. For this to work, the mercury must be an incompressible fluid so that the mass of mercury in the column is proportional to the height of the column.

The following video from the North Carolina School of Science and Mathematics (NCSSM) shows how to make a mercury barometer.

YouTube Video

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Video 9.1.1 1'35" YouTube from NCSSM on how to make a barometer.

Test Yourself

Homework: Section 9.1

Exercise 9.1.1

Why do you think the column of mercury is 744 mm Hg in the above youtube from NCSSM?

Answer available at the end of this section.

Answers

Obviously NCSSM is above sea level, but on closer look, it is in Durham NC, which is only 400 feet above sea level, and so it was also probably an arid day.

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9.2: Gas Laws

Gas Phase Chemistry

Gasses are compressible fluids which form homogenous solutions. An ideal gas is a gas where there are no attractive (or repulsive) forces between the particles and their kinetic energy (the energy of motion) is linearly proportional to the absolute temperature (K). This relationship is expressed by the Ideal Gas Law as presented in the last section.

$$\begin{aligned} PV &= nRT \\ variables &= P, V, n, T \\ constant = R &= 0.082057 \left(\frac{L \cdot atm}{mol \cdot K} \right) \end{aligned} \tag{9.2.1}$$

The value of the Ideal Gas constant R depends on the units. The bottom left unit of table $9.2.1$ (R=8.314 $\left(\frac{J}{mol \cdot K}\right)$) is of special significance and will be used extensively in the second semester of this class (when you will need to memorize that value), as it relates the energy (Joules) of a gas to the absolute Temperature (K) and number of gas particles in the gas (n). In this class you only need to memorize the first value to four significant digits, (R=0.08206 $\left(\frac{L \cdot atm}{mol \cdot K}\right)$). And hopefully you have realized by now that all these constants can be converted to SI base units (bottom right value of table, which is expressed in units of kg, m, mol and second).

Table $9.2.1$*: Select Values of R.*

R=0.082057 $\left(\frac{L \cdot atm}{mol \cdot K}\right)$	R=62.36367 $\left(\frac{L \cdot torr}{mol \cdot K}\right)$
R=8.3144598 $\left(\frac{J}{mol \cdot K}\right)$	R=8.3144598 $\left(\frac{kg(m^2)}{(s^2)mol \cdot K}\right)$

In the next section we will develop the empirical gas laws. What is important is that you be able to derive these from the ideal gas law.

Empirical Gas Laws

Empirical data is experimental data, and an experiment typically involves two variables, the independent variable, which we change, and the dependent variable, whose value changes as the independent variable is changed (section 2.10). We generically define these variables as (X,Y) where X is the independent and Y is the dependent variable, and then set up an equation where Y is a function of X, or Y =fct(X). A common function is the linear function as expressed in fig. 10.2.1

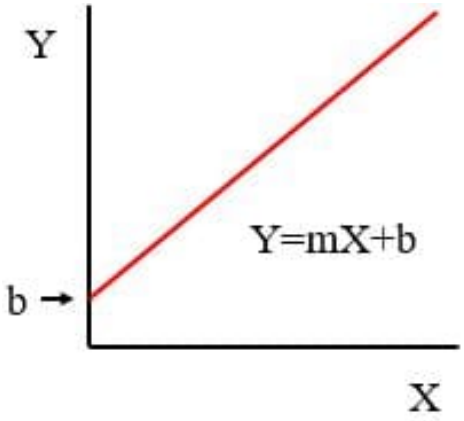


Figure $9.2.1$: Linear function where the value of Y (dependent variable) depends on the value of X (independent variable).

If we take a closer look at the ideal gas law, we see there are four variables, (P, T, V and n) and 1 constant (R). To make a plot of n

variables you need n-dimensions. So two variables gives a line, three gives a surface (x,y,x axes), and there is no way we can draw 4 variables. What the early scientists could do is hold 2 of the 4 variables constant, and make 2D plots, and the empirical gas laws are the plots that result in linear equations. These were historically determined, and then later the Ideal gas law was developed. But we can use each of the empirical gas laws as a special case of the ideal gas law, defined by which variables are constant, and which are measured. So the idea is simple, you hold two variables constant, change the third (independent variable) and observe how the fourth adjusts (dependent variable).

The Ideal Gas Law can be written as

$$\frac{PV}{nT} = R \tag{9.2.2}$$

Each of the Empirical Gas laws is an expression of the above equation where two of the four variables are held constant, and thus each of the empirical gas constants is related to the Ideal Gas constant through the values of the variables that were held constant.

Boyle's Law (constant n,T)

PV=Constant

This describes the relationship between the pressure and volume of a gas with a constant number of particles at constant temperature.

From the Ideal Gas Law we would predict:

$$PV = nRT \Rightarrow V = (nRT)\,\frac{1}{P} \tag{9.2.3}$$

since n, R and T are all constants, we have

$$PV = k \tag{9.2.4}$$

$$PV = k \quad where, k = (nRT) \tag{9.2.5}$$

Note, there are two possible graphs that come out of this relationship,

1. P vs. 1/V
2. V vs. 1/P

These graphs of of the form Y=function of X, where Y is the dependent variable and x is the independent variable. That is, if you change X, how does Y change? So if I have a sealed cylinder and I double the volume the pressure will change and so one would plot pressure as a function of volume. But if you had a frictionless piston in a cylinder and then changed the pressure in the cylinder, the piston would move until a new equilibrium was achieved, and we would say the volume is a function of pressure.

Part (a) of Figure 10.2.2 shows that P and V are not linear, but as predicted above and shown in part (b), V is linear to the reciprocal of P.

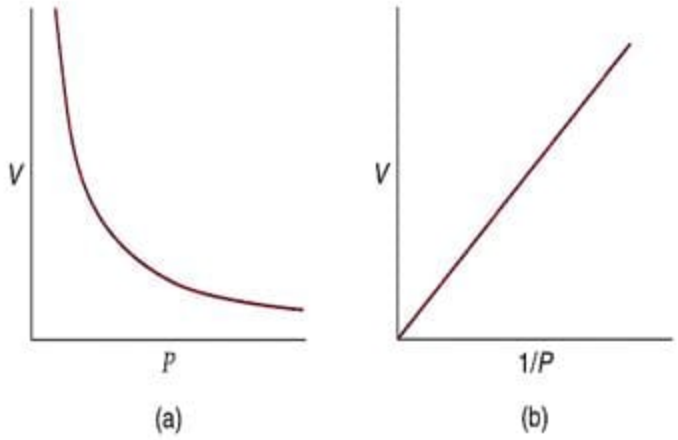


Figure 9.2.2: *Plot (a) shows a non linear relationship of V as a function of P, while (b) shows V and P are inversely related.*

What the data says is that as P goes down, 1/P goes up, and V goes up proportional to 1/P.

> Note
>
> This might be easier to understand if we switched the axes and look at how the pressure is dependent on the volume.
>
> $$P = k' \frac{1}{V} \text{ where, k=} \frac{1}{k'} \tag{9.2.6}$$
>
> The pressure is the force per unit area, and the pressure is due to the collision of gas particles against the surface of the container. The greater the collision frequency of particles with the wall, the greater the pressure. If you have a gas in a rigid container and reduce its volume, 1/V goes up, and the pressure goes up because there is a shorter path distance between the walls (increasing the collision frequency).
>
> We will revisit this after we cover the Kinetic Molecular Theory of gases, where you will find a better explanation (section 10.5.4).

Gay-Lussac's Law (constant n,V)

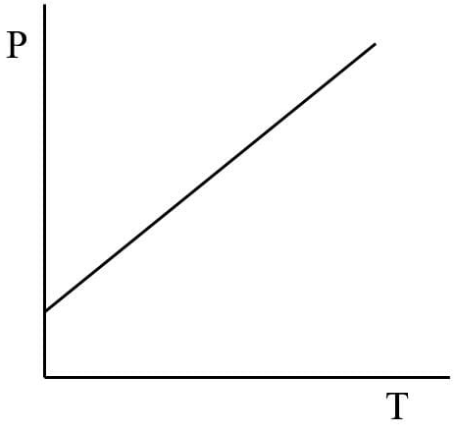
$\frac{P}{T} = Constant$

This describes the relationship between the pressure and temperature of a gas with a constant number of particles at constant temperature. From the Ideal Gas Law we would predict:

PV=nRT $\Rightarrow P = \left(\frac{nR}{V}\right) T$

since n,V & R are constants

P=kT where, k=$\left(\frac{nR}{V}\right)$



This describes a rigid closed container containing a fixed number of gas particles. As the temperature goes up the particles move faster, hit the wall more frequently at higher velocities and so the pressure goes up (remember P=$\frac{F}{A}$. As the temperature rises the pressure continues to rise and at some point the material of the rigid container will rupture and an explosion will occur as the gas escapes.

Likewise, as the temperature goes down the molecules slow down, hit the wall less frequently at lower velocities and so the pressure drops. Note, there is a lower limit to this equations, which is defined as the boiling point, and at temperatures below that Temperature the particle condenses out as a liquid (it is no longer a gas), and the Ideal gas law does not describe the system.

Charles's Law (Constant n,P)

This describes the relationship between the volume and the temperature of a gas with a constant number of particles at constant pressure. This relationship is demonstrated in the following YouTube where a balloon full of air is dropped into liquid nitrogen, and then returned to room temperature.

YouTube Video

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Video 9.2.1: 2'39" YouTube by Jefferson Lab showing the temperature effect on a balloon full of air (https://youtu.be/ZgTTUuJZAFs).

Mathematically, Charles's Law can be derived from the ideal gas equation at constant pressure and number of particles and would be modeled by an ideal balloon (closed system) that had no resistance to expansion or contraction (constant pressure).

$$PV = nRT \Rightarrow V = \left(\frac{nR}{P}\right) T \Rightarrow V = kT \; where, k = \left(\frac{nR}{P}\right) \tag{9.2.7}$$

Note, just as for Gay-Lussac's Law there is a lower limit for the temperature, being the "boiling point", at which temperature the gas converts to a liquid and no longer follows the gas laws. Figure 9.2.4 shows experimental (empirical) data for the volume of a gas as a function of temperature.

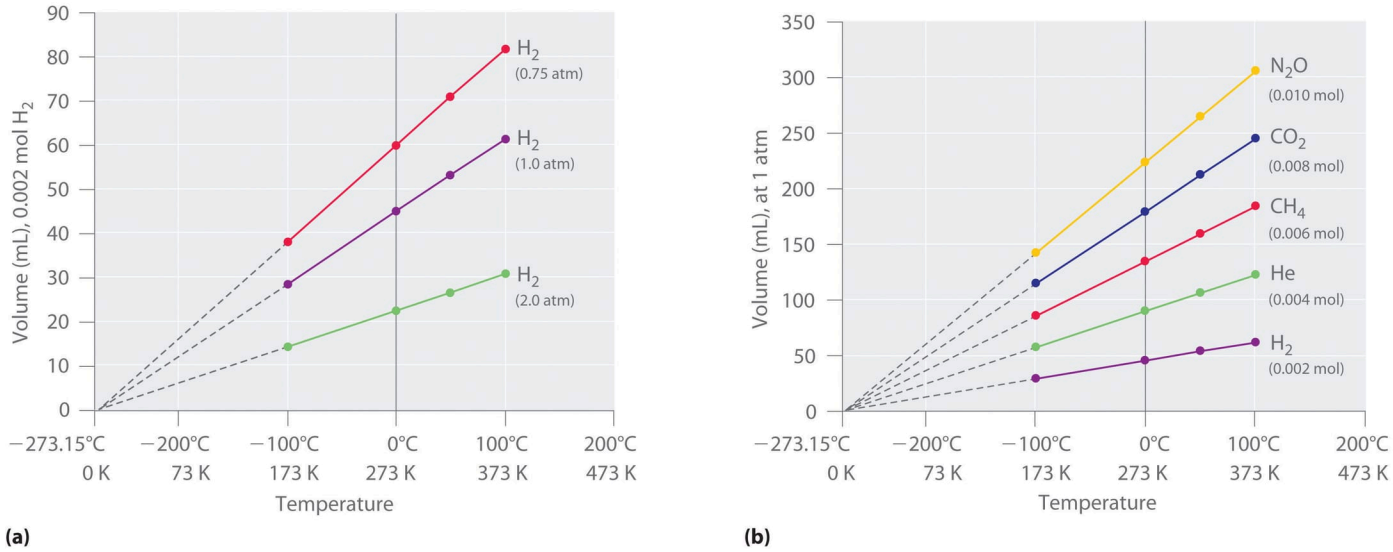


Figure 9.2.4: The Relationship between Volume and Temperature. (a) In these plots of volume versus temperature for equal-sized samples of H_2 at three different pressures, the solid lines show the experimentally measured data down to –100°C, and the broken lines show the extrapolation of the data to V = 0. The temperature scale is given in both degrees Celsius and kelvins. Although the slopes of the lines decrease with increasing pressure, all of the lines extrapolate to the same temperature at V = 0 (–273.15°C = 0 K). (b) In these plots of volume versus temperature for different amounts of selected gases at 1 atm pressure, all the plots extrapolate to a value of V = 0 at –273.15°C, regardless of the identity or the amount of the gas (CC BY-SA-NC; anonymous by request).

The significance of the invariant T intercept in plots of V versus T was recognized in 1848 by the British physicist William Thomson (1824–1907), later named Lord Kelvin. He postulated that –273.15°C was the lowest possible temperature that could theoretically be achieved, for which he coined the term absolute zero (0 K).

Avogadro's Law (Constant P,T)

This describes the relationship between the volume and the number of particles in a gas at constant pressure and temperature. We can derive Avogadro's Law from the ideal gas law using these constants:

$$\begin{aligned} PV &= nRT, \qquad \text{at constant P,T} \\ V &= \left(\frac{RT}{P}\right) n, \qquad where\ k = \left(\frac{RT}{P}\right) \\ V &= kn \end{aligned} \tag{9.2.8}$$

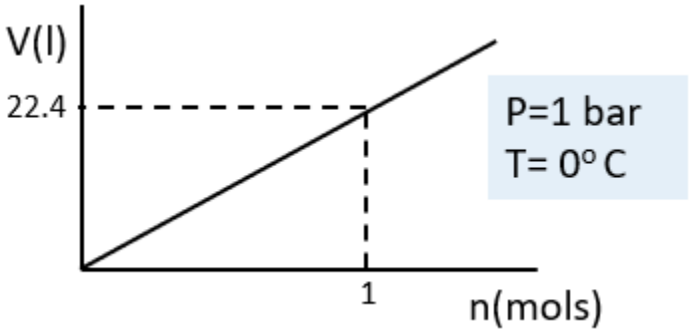


Figure 9.2.5: Avogadro's Law states that at constant P and T the volume is proportional to the number of gas particles.

An important facet of Avodadro's Law is that this relationship does not depend on the identity of the molecule (we will learn later that there are deviations), but the value of the constant does depend on the value of P and T held constant

Historically, around 1811 Avogadro postulated his hypothesis, which was that equal volumes of gasses contain equal numbers of particles, and Avogadro's Law derives from that relationship. It is common to use STP (standard temperature and pressure) as the value for the constants (T and P) in Avagadro's Law

STP (Standard Temperature and Pressure)

STP is the standard temperature and pressure that many gas values are tabulated at. You may notice different values in different textbooks because in 1982 the definiton of the standard pressure changed from 1 atm to 1 bar.

- Until 1982, STP was defined as a T= 0 °C and P = 1 atm (101.325 kPa).
- Since 1982, STP is defined as a T= 0 °C and P = 1 bar (100kPa)

Molar Volume at STP

Avogadro's hypothesis states that the volume of a mole should be the same for any substance, and as can be seen from Table 9.2.1.

Table 9.2.1: Molar Volumes of Several Gases at STP

Substance	Formula	Molar Volume/liter mol^{-1}
Hydrogen	$H_2(g)$	22.43
Neon	$Ne(g)$	22.44
Oxygen	$O_2(g)$	22.39
Nitrogen	$N_2(g)$	22.40
Carbon dioxide	$CO_2(g)$	22.26
Ammonia	$NH_3(g)$	22.09

The data in Table 9.2.1, then, indicate that for a variety of gases, 6.022×10^{23} molecules occupy almost exactly the same volume (the ***molar volume***) if the temperature and pressure are held constant, and the molar volume for an ideal gas is 22.4 l/mol. So if you can measure the volume of a gas at STP, you can calculate the number of moles.

Ideal Gas Law Calculations

There are two basic types of calculations we will do involving the ideal gas laws.

1. PV=nRT calculations.
2. Two State Problems

The two state problem will be approached in the next secion.

PV=nRT

These are simple algebraic problems where you are given 3 of the 4 variables. The trick is to keep your units in your equations and be sure the units of the answer make sense. If you think about it, the numerical value of R depends on the units of R (see table 10.2.1), and your variables must have the units of R or they will not cancel. This is especially important for temperature, which must be in units of Kelvin.

Exercise 9.2.1

The following problems represent various solutions to the ideal gas equation

a. A sample of gaseous Cl_2 has a volume of 12.4 L at 500.0 K and 0.521 atm. How many moles are present?
b. What is the mass of chlorine gas in a 12.4L container at 500.0K and 0.521 atm?
c. What volume is occupied by 1.06 mol of CO_2 gas at 299 K and a pressure of 0.89 atm?
d. What is the temperature of 1.41 mol of methane gas in a 5.0 L container at 1.00 atm?
e. What is the pressure exerted by a 1.75 mol sample of water at 7.0L and 20°C?

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

- Bob Belford (UALR) and November Palmer (UALR)
- Adoptions from Openstax

Answers

$$n = \frac{PV}{RT} = \frac{0.521atm\,(12.4L)}{0.08206\frac{L\cdot atm}{mol\cdot K}\,(500K)} = 0.157mol$$

$$n = \frac{PV}{RT} = \frac{0.521atm\,(12.4L)}{0.08206\frac{L\cdot atm}{mol\cdot K}\,(500K)} = 0.157mol$$

$$0.157mol\ Cl_2\left(\frac{70.90g\ Cl_2}{mol}\right) = 11.1g \tag{9.2.9}$$

$$V = \frac{nRT}{P} = \frac{1.06mol\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)299K}{0.89atm} = 29L$$

$$T = \frac{PV}{nR} = \frac{1.00atm\,(5.0L)}{1.41mol\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)} = 43K$$

$$P = \frac{nRT}{V} = \frac{1.75mol\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)293.15K}{7.0L} = 6.0atm$$

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9.3: Two State Ideal Gas Problems

Two State Approach

If you have a gaseous system and then perturb it by changing one of more variables the other variables adjust to that change. The two state approach is a quick way to solve these problems. The idea is simple, you have a relationship PV=nRT, so at one state, $P_1V_1 = n_1RT_1$ and $P_2V_2 = n_2RT_2$, so we can set up the equivalence:

$$\frac{P_1V_1}{P_2V_2} = \frac{n_1RT_1}{n_2RT_2} = \frac{n_1T_1}{n_2T_2} \tag{9.3.1}$$

so

$$\frac{P_1V_1}{P_2V_2} = \frac{n_1T_1}{n_2T_2} \tag{9.3.2}$$

(note: the constant (R) cancels out)

Warning: T must be in Kelvin, as the equation PV=nRT requires T to be in Kelvin, and even though R cancels in the two state approach, T still needs to be in Kelvin.

When solving these problems it is wise to make a table and write down what you are given. If something is held constant, it cancels out of the equation.

State 1	State 2
T_1=	T_2=
P_1=	P_2=
V_1=	V_2=
n_1=	n_2=

Example $9.3.3$

What is the volume of a perfectly elastic balloon if it rises to an altitude where the pressure drops to 0.100 atm and the temperature drops to 173 K if it had a volume of 5.00 liters at sea level where the pressure was 1.00 atm and the temperature was 273K?

Solution

(a) Set up two state equation and solve for V_1 or V_2 (which either is easier, as we have not assigned the data to the states

$$\frac{P_1V_1}{P_2V_2} = \frac{n_1T_1}{n_2T_2} \tag{9.3.3}$$

$$becomes$$

$$V_1 = V_2\left(\frac{n_1}{n_2}\right)\left(\frac{T_1}{T_2}\right)\left(\frac{P_2}{P_1}\right)$$

(b) Fill our the data table so we properly identify the data with the state.

TIP, Since we want the volume in the upper atmosphere we will call that state 1, but we could have called sea level state one, in which case we would solve for V_2.

State 1	State 2
T_1= 173K	T_2= 273K
P_1= 0.100 atm	P_2= 1.00 atm
V_1= ?	V_2= 5.00 L
n_1=	n_2=

NOTE: This is a closed system and no gas entered or escaped from the balloon and so n_1=n2 and they cancel.

$$V_1 = 5.00L \cancel{\left(\frac{n_1}{n_2}\right)} \underbrace{\left(\frac{173K}{273K}\right)}_{\substack{T \text{ must be} \\ \text{in Kelvin}}} \left(\frac{1atm}{0.100atm}\right) \tag{9.3.4}$$

The solution to this problem is demonstrated in video 9.3.4.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 9.3.4: 5'10" YouTube solving this problem (https://youtu.be/BCYEZv9IQFA)

Note

Many textbooks cover the "combined gas law":

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \tag{9.3.5}$$

On closer inspection the combined gas law is applying the two-state approach to the ideal gas law where n, the number of molecules is constant, and that was the case we just solved (example 10.3.2). The advantage of the two state approach is that it not only allows us to handle changes in the number of particles, but it is a technique that can be applied to many equations, not just the ideal gas equation.

If you have a relationship y=mx, you can always say that $\frac{y_1}{y_2} = \frac{x_2}{x_2}$, and during the second semester this technique will be applied to power functions ($Y=AX^m$) and exponential functions $Y=Ae^{xm}$. It is worth mastering the two state approach now.

Test Yourself

Homework: Section 9.3

Graded Assignment: Section 9.3

Exercise 9.3.2

Solve the following problems using the two state approach.

a. What pressure would 6.1 moles of a gas in a rigid container have at 20.0 °C have if it had a pressure of 0.45 atm at -45.0 °C?
b. A pressure tank containing chlorine gas at a pressure of 2.00 atm and a temperature of 40°C is set with a pressure relief valve set to open at a pressure of 10.0 atm. At what temperature will the relief valve open?
c. A gas is in a sealed cylinder at 25°C with a piston which can expand or contract to change the volume. The initial volume and pressure are 75.0 L and 980.0 torr. A force is applied to the piston and the volume adjusts to a final pressure and temperature of 5.00 atm at 188°C. What is the final volume?

Answer a available at the end of this section.

Answers

$$P_1 = P_2\frac{T_1}{T_2} = (0.45atm)\left(\frac{293.15K}{228.15K}\right) = 0.58atm \tag{9.3.6}$$

$$T_2 = T_1\frac{P_2}{P_1} = (313.15K)\left(\frac{10.0atm}{2.00atm}\right) = 1570K \tag{9.3.7}$$

$$V_2 = V_1\left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right) = 75.0L\left(\frac{980torr\left(\frac{1atm}{760torr}\right)}{5.00atm}\right)\left(\frac{461.15K}{298K}\right) = 29.9L \tag{9.3.8}$$

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9.4: Gas Phase Stoichiometry

Introduction

Now that we have covered the ideal gas law we can revisit the stoichiometric equations of Chapter 4 and account for gas phase reactants and products in a more meaningful manner. While going over this material you may wish to review the following sections.

- 3.2: Types of Chemical Reactions (combustion and many types of single displacement involve gasses)
- 3.5.4: Acid Reactions Evolving Gases (Strong acid reactions with salts of oxyanions [acid anhydrides])
- 4.1: Stoichiometry
- 4.2: Limiting & Excess Reagents

In the previous sections we would report gasses in terms of their moles, or even their mass. Lets take a second look at Exercise 4.2.1, where we calculated the mass of carbon dioxide formed upon the combustion of 25.00g glucose with 40.0 g of oxygen. Sure we could solve the problem, but at standard conditions oxygen or carbon dioxide are gasses not solids, and so you would not be able to directly measure their mass, and a more realistic problem would be to define them in terms of measurable quantities like pressure, temperature and volume.

Overview of Stoichiometry:

In stoichiometric calculations you typically have three steps.

1. Convert measured quantities to moles (assuming you don't start with moles, but something measurable)
2. Use balanced equations to relate moles of starting material to moles of desired substance sought.
3. Convert moles of desired substance to measurable quantities (assuming you don't want moles, but something measurable)

mass	solute in solution	
n=$\frac{m}{fw}$	n=MV	n=$\frac{PV}{RT}$
m=mass fw=formula weight	M=Molarity=$\frac{mol_{solute}}{liter_{solution}}$ V=Volume in liters	P=Pressure V=Volume R=Ideal Gas Constant T=Temperature in Kelvin

As in Chapter 4, it is a good idea to write out and balance the equation, write under it what is given, and then solve the problem. The following problems involves calculations involving all three phases.

Example $9.4.1$

What volume of gas would be produced at 760.0 torr and 298 K if 1.00 g of magnesium metal is dropped into 50.0 ml of 1.00 M Hydrochloric Acid?

Solution
(See Video $9.4.1$ for solution)

Step 1: Write out and balance equation and write underneath it what is given:

$$\underset{\substack{m=1.00g\\ fw=24.3\frac{g}{mol}}}{Mg(s)} + \underset{\substack{0.0500L\\ 1.00M}}{2HCl(aq)} \rightarrow Mg(Cl)_2(aq) + \underset{\substack{P=760.0atm\\ T=298K\\ n=?\\ V=??}}{H_2(g)} \quad (9.4.1)$$

NOTE: We want V_{H_2}, but have two unknowns (V and n), so we need to calculate n

Step 2: Calculate theoretical yield of n (based on complete consumption of limiting reagent)

$$1.00g\ Mg\left(\frac{mol\ Mg}{24.3g}\right)\left(\frac{1}{1mol\ Mg}\right)=0.0411 \tag{9.4.2}$$
$$0.0500L\left(\frac{1.00mol\ HCl}{L}\right)\left(\frac{1}{2mol\ HCl}\right)=.025$$
$$HCl=limiting\ reagent$$
$$n_{H_2}=0.025\left(1mol\ H_2(g)\right)=0.025mol\ H_2(g)$$

Step 3: Now that we have n_{H_2}, use ideal gas law to solve for V at the conditions of the problem.

$$V=\frac{nRT}{P}=\frac{0.02500mol\left(0.08206\left(\frac{l\cdot atm}{mol\cdot K}\right)\right)298K}{760\ torr\left(\frac{1atm}{760\ torr}\right)}=.6113L=611mL \tag{9.4.3}$$

YouTube Video

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Video 9.4.1: 5'35" YouTube showing solution for problem above

Test Yourself

Homework: Section 9.4

Graded Assignment: Section 9.4

Exercise 9.4.1

What volume of oxygen is formed at 750.0torr and 25.0°C if 1.30 g potassium chlorate decomposes into potassium chloride and oxygen?

Answer available at the end of this section.

Exercise 9.4.2

Solve the following problems

a. 5.00 g of Mg is added to 50.0 mL of 0.800M HCl. A single displacement reaction occurs in a sealed container at 25.0°C. What is the pressure of the hydrogen gas generated if the volume above the solution is 100.0 ml, and we ignore any pressure due to the evaporated water?
b. Nitrogen dioxide is formed in a closed rigid container at 90°C and 1.00 atm when 1.50 g NO and 2.00 mole of O_2 are mixed. After the reaction is finished the pressure changes to 3.90 atm. What is the final temperature?

Answer a available at the end of this section.

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Answers

Exercise 9.4.1

First write out balanced equation and calculate moles oxygen formed

$$\underset{\substack{m=1.30g\\ fw=122.55\frac{g}{mol}}}{2KClO_3(s)} \rightarrow 2KCl(s) + \underset{\substack{P=750torr\\ T=298K\\ n=?\\ V=??}}{3O_2(g)} \tag{9.4.4}$$

First calculate moles oxygen produced by the complete decomposition of the potassium chlorate

$$1.30gKClO_3\left(\frac{mol\ KClO_3}{122.55g}\right)\left(\frac{3mol\ O_2}{2mol\ KClO_3}\right) = 0.0159mol\ O_2 \tag{9.4.5}$$

Substitute into ideal gas law and solve

$$V = \frac{nRT}{P} = \frac{0.0159mol\left(0.08206\frac{l\cdot atm}{mol\cdot K}\right)298K}{750torr\left(\frac{1atm}{760torr}\right)} = 394mL \tag{9.4.6}$$

First you need to calculate the theoretical yield based on the complete consumption of the limiting reagent.

$$5g\ Mg\left(\frac{mol\ Mg}{24.31g}\right)\frac{1}{1molMg} = 0.206 \tag{9.4.7}$$
$$0.0500L\left(\frac{0.800\ mol\ HCl}{L}\right)\frac{1}{2\ mol\ HCl} = 0.0200$$

so HCl is the limiting reagent and the theoretical yield is

$$0.0500L\left(\frac{0.800\ mol\ HCl}{L}\right)\frac{1\ mol\ H_2}{2\ mol\ HCl} = 0.0200\ mol\ H_2 \tag{9.4.8}$$

Substituting into the ideal has equation gives the pressure

$$P_{H_2} = \frac{n_{H_2}RT}{V} = \frac{0.0200mol\ H_2\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)298K}{0.100L} = 4.89atm \tag{9.4.9}$$

This combines stoichiometry with the two state approach. First write out and balance the equation, indicating initial reactants

$$\underset{\substack{m=1.50g\\ fw=30.01\frac{g}{mol}}}{2NO(g)} + \underset{2.00\ mol}{O_2(g)} \rightarrow \underset{\substack{P=1.00atm\\ T=373.15K\\ n=?\\ V=??}}{2NO_2(g)} \tag{9.4.10}$$

Now set up your two state table:

	Initial State (1)	Final State (2)
Temp	90+273.15=373.15K	T_2=?
Pressure	1.00 atm	3.90 atm
moles (n)	n_1=	n_2=
Vol	V_1=V_2	V_1=V_2

First calculate n_1, the initial number of moles of all species (before reaction):

$$n_1 = n_{NO} + n_{O_2} + n_{NO_2} = 1.50g\left(\frac{mol\ NO}{30.01g}\right) + 2.00mol\ O_2 + 0.00mol\ NO_2 \tag{9.4.11}$$
$$= 0.04998mol\ NO + 2.00mol\ O_2 = 2.05mol_{total}$$

Second calculate n_2, the total number of moles all species after reaction

NO is the limiting reagent because

$$\frac{0.05}{2} < \frac{2.00}{1} \tag{9.4.12}$$

the amount of NO_2 produced is:

$$0.05mol\ NO\left(\frac{2mol\ NO_2}{2mol\ NO}\right) = 0.05mol\ NO_2 \tag{9.4.13}$$

the amount of excess oxygen is

$$O_2(excess) = O_2(initial) - O_2(consumed) \tag{9.4.14}$$
$$= 2.00mol - 0.05mol\ NO\left(\frac{mol\ O_2}{2mol\ NO}\right) = 2.00 - 0.025 = 1.975$$

total moles after reaction

$$n_2 = n_{NO} + n_{O_2} + n_{NO_2} = 0 + 1.975 + 0.05 = 2.025 \tag{9.4.15}$$

$$T_2 = T_1 \left(\frac{n_1}{n_2}\right)\left(\frac{P_2}{P_1}\right) = 363.15K \left(\frac{2.05mol}{2.025mol}\right)\left(\frac{3.90atm}{1.00atm}\right) = 1433.77K = 1400K \tag{9.4.16}$$

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10: THERMODYNAMICS



- Topic hierarchy

10.1: Energy

10.2: Heat Capacity

10.3: Energy and Phase Changes

10.4: First Law of Thermodynamics

10.5: Enthalpy of Reaction

10.6: Enthalpy Calculations

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10.1: Energy

What is Energy?

Energy is one of the most fundamental and universal concepts of physical science, but one that is remarkably difficult to define in a way that is meaningful to most people. This perhaps reflects the fact that energy is not a “thing” that exists by itself, but is rather an attribute of matter (and also of electromagnetic radiation) that can manifest itself in different ways. It can be observed and measured only indirectly through its effects on matter that acquires, loses, or possesses it.

The concept that we call energy was very slow to develop; it took more than a hundred years just to get people to agree on the definitions of many of the terms we use to describe energy and the interconversion between its various forms. But even now, most people have some difficulty in explaining what it is; somehow, the definition we all learned in elementary science ("the capacity to do work") seems less than adequate to convey its meaning.

Kinetic energy and potential energy

Kinetic energy is associated with the **motion** of an object, and its direct consequences are part of everyone's daily experience; the faster the ball you catch in your hand, or the heavier it is, the more you feel it.

Potential energy is energy a body has by virtue of its *location*. But there is more: the body must be subject to a "restoring force" of some kind that tends to move it to a location of lower potential energy. Think of an arrow that is subjected to the force from a stretched bowstring; the more tightly the arrow is pulled back against the string, the more potential energy it has.

More generally, the restoring force comes from what we call a *force field*— typically a gravitational, electrostatic, or magnetic field. We observe the consequences of *gravitational* potential energy all the time, such as when we walk, but seldom give it any thought.

"Chemical energy"

Electrostatic potential energy plays a major role in chemistry; the potential energies of electrons in the force field created by atomic nuclei lie at the heart of the chemical behavior of atoms and molecules. "Chemical energy" usually refers to the energy that is stored in the chemical bonds of molecules. These bonds form when electrons are able to respond to the force fields created by two or more atomic nuclei, so they can be regarded as manifestations of electrostatic potential energy.

Interconversion of potential and kinetic energy

Transitions between potential and kinetic energy are such an intimate part of our daily lives that we hardly give them a thought. It happens in walking as the body moves up and down. Our bodies utilize the chemical energy in glucose to keep us warm and to move our muscles. In fact, life itself depends on the conversion of chemical energy to other forms.

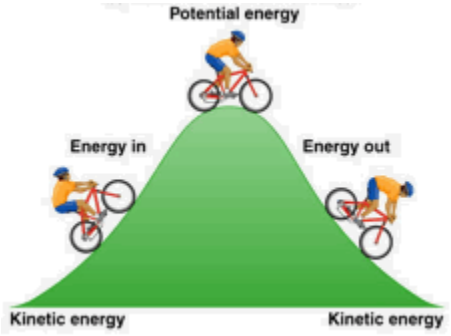


Figure: Conservation of energy applied to a bicyclist and a hill.

Energy is conserved: it can neither be created nor destroyed. So when you go uphill, your kinetic energy is transformed into

potential energy, which gets changed back into kinetic energy as you coast down the other side. And where did the kinetic energy you expended in peddling uphill come from? By conversion of some of the chemical potential energy in your breakfast cereal.

Thermal energy

Kinetic energy is associated with motion, but in two different ways. For a macroscopic object such as a book or a ball, or a parcel of flowing water, it is simply given by $\frac{1}{2}mv^2$. However, when an object is dropped onto the floor, or when an exothermic chemical reaction heats surrounding matter, the kinetic energy gets dispersed into the molecular units in the environment. This "microscopic" form of kinetic energy, unlike that of a speeding bullet, is completely random in the kinds of motions it exhibits and in its direction. We refer to this as "thermalized" kinetic energy, or more commonly simply as *thermal energy*. We observe the effects of this as a rise in the temperature of the surroundings, or a phase change (like melting). The temperature of a body is direct measure of the quantity of thermal energy is contains.

Energy units

Work based definition of energy (Joule)

This is the SI unit of energy and one Joule is a Newton meter (J = N·m), where a Newton is the SI unit of force. The Newton (named after Isaac Newton) comes from classical mechanics and Newton's second Law of motion, F = ma (mass times acceleration), and acceleration is the change in velocity per unit time, $a = \frac{\Delta v}{\Delta t}$ and velocity is the vectorial change in displacement per unit time, $v = \frac{\Delta x}{\Delta t}$, where x is the displacement in a specific direction. So like any derived SI unit, the Joule can be described in terms of the SI base units, $1J = 1kg\frac{m^2}{sec^2}$

Heat based definition of energy (Calorie)

Historically, the calorie is the energy required to raise one gram of water one degree celsius (from 14.5°C to 15.5°C), but today the calorie is considered to be a depreciated unit, and according to NIST Special Publication 1038, "*The International System of Units (SI)-Conversion Factors of General Use*", the definition of the calorie is based on the Joule:

1cal ≡ 4.184 Joules

Note, the above is a defined number and should not be treated as if it has significant digits. Also, the calorie is actually a very small unit of measurement and it should be noted that **the nutritional Calorie (written with a capital C) is really a kilocalorie**.

Heat and Work

Heat and work are both measured in energy units, so they must both represent energy. How do they differ from each other, and from just plain “energy” itself? In our daily language, we often say that "this object contains a lot of heat", but this is gibberish in thermodynamics terms, although it is ok to say that the object is "hot", indicating that its temperature is high. The term "heat" has a special meaning in thermodynamics: it is a *process* in which a body (the contents of a tea kettle, for example) acquires or loses energy as a direct consequence of its having a *different temperature* than its surroundings. Hence, thermal energy can only flow from a higher temperature to a lower temperature. It is this flow that constitutes "heat". Use of the term "flow" of heat recalls the incorrect 18th-century notion that heat is an actual substance called “caloric” that could flow like a liquid.

Test Yourself

Query 10.1.1

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Query 10.1.2

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- Liliane Poirot

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10.2: Heat Capacity

Learning Objectives

- Define heat capacity and specific heat capacity and differentiate between the two terms
- Deduce which substance will have greatest temperature changed based on specific heat capacities
- Calculate unknown variables based on known variables using the specific heat equation

Heat Capacitance

The heat capacity of a substance describes how its temperature changes as it absorbs or releases heat, it is the capacity of a substance to contain heat. This equation relates the heat absorbed (or lost) to the temperature change

$$\Delta q = q_{\text{transferred}} = q = C\Delta T \tag{10.2.1}$$

Please note this heat q is the amount of heat transferred to or from an object as its temperature is changed, and we call it q, not (\Delta q\). It is not the total heat energy of the object.

Rearranging the above Equation gives the definition of the heat capacitance of an object

$$C = \frac{q(J)}{\Delta T(K)} \tag{10.2.2}$$

This is an important equation because it relates the heat lost or gained by an object to its temperature change. A substance with a small heat capacitance cannot hold a lot of heat energy and so warms up quickly. We should note that the rate of heat transfer (how fast something heats up or cools down), is a function of the temperature difference. So as heat transfers from the hot object (which cools) to the cold object (which warms) the temperature difference reduces, and the rate of heat transfer slows down and finally stops when they reach the same temperature. The heat capacitance of a substance depends both on the material it is made of, and the mass of the substance.

Note: You can determine the above equation from the units of Capacity (energy/temperature). That is if a constant has units, the variables must fit together in an equation that results in the same units. So C equals something with energy in the numerator and temperature in the denominator. Now, you need to use some common sense here, as we are adding heat, not work, and adding heat changes the temperature, it does not make the temperature. So the right side is a ΔT, and not a T.

Exercise 10.2.1

The Heat Capacitance of a 10.0 g silver ring is 2.36J/^{o}C.

a. Convert this value to units of J/K.
b. Convert this value to units of cal/^{o}C.

Answer a available at the end of this section.

Specific Heat Capacitance

The specific heat capacitance is the heat capacitance per gram of a substance. This value depends on the nature of the chemical bonds in the substance, and its phase.

$$q = mc\Delta T \tag{10.2.3}$$

or

$$c = \frac{q(J)}{m(g)\Delta T(K)} \tag{10.2.4}$$

Note: Capital "C" is the **Heat Capacity** of an object, lower case "c" is the **specific heat capacity** of a substance. The heat capacity of an object made of a pure substance is:

$$C = mc \tag{10.2.5}$$

If the material of an object is made of uniform in composition you can use the specific heat capacity for that material to calculate the heat capacitance of the object. So doubling the mass of an object doubles its heat capacity, but does not change its specific heat capacitance.

Exercise 10.2.2

In exercise 10.2.1 we saw that a 10.0 g silver ring has a heat capacitance of 2.36J/^{o}C, what is the specific heat capacity of silver?

Answer available at the end of this section.

Tip: Using Units of a Constant to the Identify equation

It should be noted that just as for heat capacity, the units of specific heat capacity must align with the units of the equation, and so you can calculate the equation from the units, as long as you realize J is a unit of energy (we are talking heat, not work), g is a unit of mass, and °C is a unit of temperature, although here, it stand for temperature change (ΔT). **Video** 10.2.1 **shows how we can use the units of a constant to determine the equation.**

YouTube Video

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Video 10.2.1: 5'32" YouTube on using constants to determine equations related to heat capacity and phase changes (https://youtu.be/mWj9pHQOyIc). **NOTE:** in this video the values of ΔH_v are for water, and so we use the molar mass of water to relate these. Some of these values are tabulated in table 5.3.1.

Note the heat capacity depends on the phase of the substance.

Table 10.2.1: Specific Heat Capacities for common substances.

Substance	Symbol (state)	Specific Heat (J/g °C)		Substance	Symbol (state)	Specific Heat (J/g °C)
helium	He(*g*)	5.193		aluminum	Al(s)	0.897
water	$H_2O(l)$	4.184		carbon dioxide	$CO_2(g)$	0.853
ethanol	$C_2H_6O(l)$	2.376		Silicon	Si(s)	0.712
ice	H_2O(s)	2.093 (at –10 °C)		argon	Ar(g)	0.552
water vapor	$H_2O(g)$	1.864		iron	Fe(s)	0.449
nitrogen	$N_2(g)$	1.040		copper	Cu(s)	0.385
air	mixture	1.007		gold	Au(s)	0.129
oxygen	$O_2(g)$	0.918		lead	Pb(s)	0.128

Note: Metals have low heat capacities and thus undergo rapid temperature rises when heat is applied.

Exercise 10.2.3

If you add the same amount of heat to an equal mass of liquid water, solid gold, and solid iron, which would end up having the highest temperature?

Answer available at the end of this section.

Example 10.2.1

What is the final temperature if 100.0 J is added to 10.0 g of aluminum at 25°C? $c_{Al} = 0.902\ J/(g.^oC)$. For extra help watch video 10.2.2, which solves this problem.

Solution

$$T_f = T_i + \frac{q}{mc} = 25^oC + \frac{100.0\,J}{10.0\,g(\frac{0.902\,J}{g\cdot^0C})} = 36.1^oC$$

YouTube Video

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Video 10.2.2: 1'18" YouTube solving example 10.2.2 (https://youtu.be/4hKfm4B-C6o).

Specific heat capacity can be used to identify an unknown substance

The specific heat capacity is a physical property of the material a substance is composed of and can be used to help identify the substance the way density can help identify an incompressible substance like a solid or liquid. It should be noted that two substances can have the same specific heat capacity just as two substances can have the same density, but for example, if the heat capacity of a clear liquid is not 1 cal/g°C, the substance can not be pure water.

Example 10.2.2

You have an unknown metal that is either Al, Cu, Ag or Fe and want to identify it. Upon adding 51.26J to 10.0G of the metal its temperature is raised by 22°C

Solution

Calculate the specific heat capacity and relate to those on table Table 10.2.1. The metal is silver and this problem is solved in Video 10.2.3

YouTube Video

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Video 10.2.3: 1'12" YouTube using specific heat capacity calculations to help identify a metal (https://youtu.be/duAzx66-0TI)

Exercise 10.2.4

Can you use the specific heat capacity to tell the difference between lead and gold?

Answer available at the end of this section.

Test Yourself

Homework: Section 10.2

Query 10.2.1

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Query 10.2.2

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- Ronia Kattoum & Emily Choate (UALR)
- Liliane Poirot

Answers

2.36J/K, these are the same value because the denominator is $DeltaT$ not T.

0.564cal/^{o}C.

Assuming the ring is pure silver, the specific heat capacity of silver is 0.236J/g^{o}C.

Solid Gold. They all have the same mass and are exposed to the same amount of heat. So, the one with the lowest specific heat would have the highest temperature. It has the lowest resistance to temperature change when exposed to heat. If you ever reached into an oven to grab your food with a gold bracelet on, you may have experience the low specific heat capacity of gold. Metals have low heat capacities and thus undergo rapid temperature rises when heat is applied.

No, gold had a specific heat capacity of 0.129J/g^{o}C and lead of 0.128J/g^{o}C. The thousandths position is uncertain and so to three significant digits, you can not differentiate between these two samples (you report all certain and the first uncertain, so if you have a measurement of 0.128, you are not really sure of the 0.008 value.

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10.3: Energy and Phase Changes

Learning Objectives

- Define heat of fusion and heat of vaporization
- Relate heat of fusion and heat of vaporization to heat of freezing and heat of condensation
- Label various part of a heating curve
- Using a heating curve, calculate the heat input/output associated with phase changes and temperature changes using the proper equation(s)
- Using a heating curve, calculate the final temperature for a given quantity of heat input/out.

Energy and Phase Changes

When adding or removing heat from a system one of three things can occur:

1. Temperature Change within a phase (physical change)
2. Phase Change between two phases (physical change)
3. Adding heat can cause a chemical reaction (chemical change)

In this section we shall look at changes between two phases. In reality there are 4 states of matter, solid, liquid, gas and plasma, although in this class we will focus on the three common states, and their transitions.

YouTube Video

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Video \({10{.}3{.}\eatSpaces{1} \endSpaces}/): 0'49" YouTube uploaded by kosasihiskandarsjah describing the four states of matter (https://youtu.be/tEY0V8BD8K4)

Figure 10.3.1 shows the phase transitions between the various phases and not the enthalpy of the system is a measure of the heat energy at constant pressure and will be introduced in section 5.4.3 of this Chapter.

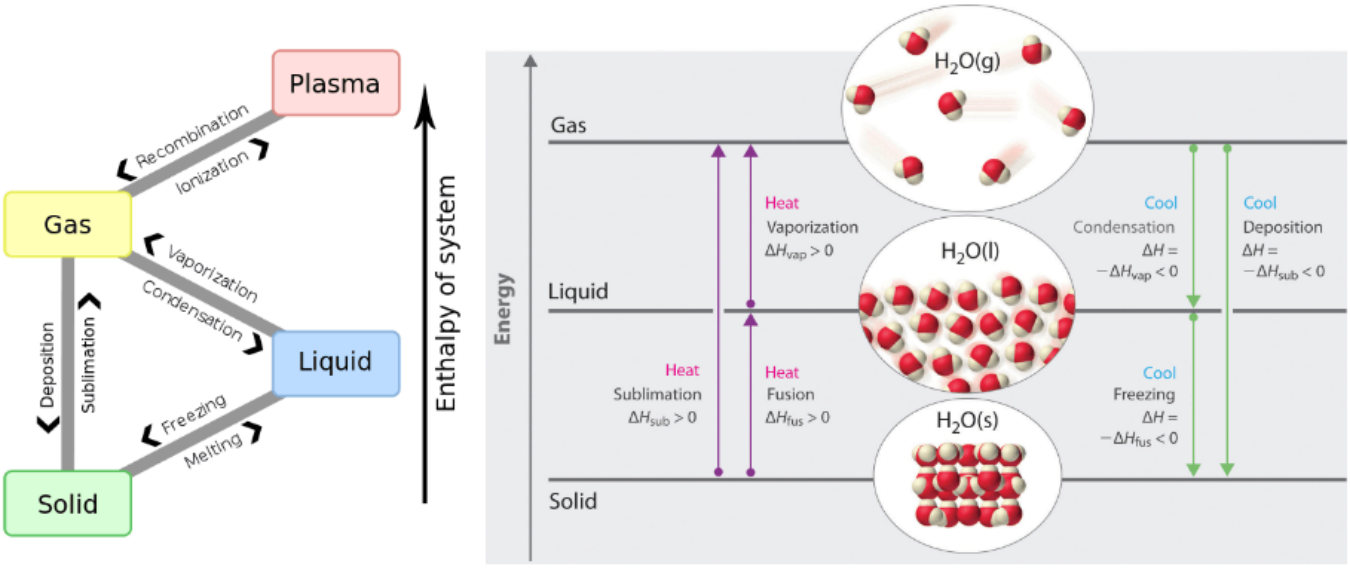


Figure 10.3.1: Energy diagrams showing phase transitions. The diagram on the left shows all 4 states, while the one on the right shows the three common states. This class will not be dealing with transitions to the plasma state, but you should be aware of it. There are several things to note on the right diagram.

1. Transitions going up (increased energy) are endothermic and require energy ,while those going down (decreased energy) are exothermic and release energy.
2. The energies absorbed in the endothermic steps are identical in magnitude to the energies released in the exothermic steps for the same transition.
   $\Delta H_{Fusion} = -\Delta H_{Freezing}$
   $\Delta H_{Vaporization} = -\Delta H_{Condensation}$
   $\Delta H_{Sublimation} = -\Delta H_{Deposition}$.
3. Energy is conserved
   $\Delta H_{Sublimation} = \Delta H_{Fusion} + \Delta H_{Vaporization}$
   $\Delta H_{Deposition} = \Delta H_{Condensation} + \Delta H_{Freezing}$

Note: Endothermic and Exothermic Reactions

We will formally introduce endothermic and exothermic reactions in the section 5.4 (the next section) when we discuss the First Law of Thermodynamics and define the direction of heat transfer (as being positive when something gains energy, and negative when something loses energy). So **an exothermic process is one where energy is released to the environment** and **an endothermic one is where energy is absorbed from the environment**. We are introducing the concept early because all the process going in direction of increased energy (S->L ->G) are endothermic (you add energy to boil water), and those of decreasing energy (G -> L -> S) are exothermic

Exercise 10.3.1

Consider the process of water freezing.

a. Is it an endothermic or exothermic process?
b. Does it cool or warm the surroundings?

Answer a available at the end of this section.

Sublimation

Students often have a hard time understanding sublimation. Carbon dioxide (dry ice) that you can buy at many grocery stores sublimes, that is, it goes directly from a solid to a gas without ever melting (which is why the call it "dry ice"). Video 10.3.2

demonstrates the sublimation and deposition of iodine.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 10.3.2: 1'39" YouTube uploaded by koen2all showing sublimation and deposition of iodine (https://youtu.be/jX9pskbKSw0).

Quantifying Phase Changes

We have seen how heat is either released or absorbed during phase changes. Now, we will study the quantity of energy associated with those phase changes. In order to carry out those calculations, we need to establish a standard bases for comparison, the **molar heat of the phase transition,** or the heat required for one mole of substance to change phase. Particularly, **molar heat of fusion,** refers to the amount of heat (in kJ) it take for one mole of substance to transition from the solid state to the liquid state (melt). The **molar heat of vaporization**, is the amount of heat (in kJ) it takes for one mole of a substance to transition from the liquid state to the vapor phase. Since we are not always working with exactly one mole of each substance, we will use this equation to Figure out the heat for the particular number of moles that we have according to the following equation:

$$\mathbf{q = n\Delta H_{transition}}$$

where n is the number of moles, and H is the heat or enthalpy required for one mole of a substance to go through the phase transition at the temperature where it occurs (which depends on the pressure). Unless otherwise noted, you can assume these are the boiling points and freezing point at 1 atm pressure. As you can see from Table 1, these values are usually given and tabulated. Since, we typically deal with the mass of substance, there are some tables that also give you the heat of transition of a substance on a per gram bases (usually reported as J/g) Here is a link the following tables heat of fusion and heat of vaporization tables:

B2: Heats of Vaporization (Reference Table)

B3: Heats of Fusion (Reference Table)

Substance	Melting Point (°C)	ΔH_{fus} (kJ/mol)	Boiling Point (°C)	ΔH_{vap} (kJ/mol)
N_2	−210.0	0.71	−195.8	5.6
HCl	−114.2	2.00	−85.1	16.2
Br_2	−7.2	10.6	58.8	30.0
CCl_4	−22.6	2.56	76.8	29.8
CH_3CH_2OH (ethanol)	−114.1	4.93	78.3	38.6
$CH_3(CH_2)_4CH_3$ (*n*-hexane)	−95.4	13.1	68.7	28.9
H_2O	0	6.01	100	40.7
Na	97.8	2.6	883	97.4
NaF	996	33.4	1704	176.1

Table 10.3.1: Enthalpies of Fusion (melting) and Vaporization for some common substances, and the transition temperature where

they occur (at normal pressure of 1 atm).

Exercise 10.3.2

For a given substance, which is higher, ΔH_{Fusion} = $\Delta H_{Vaporization}$ and why?

Answer available at the end of this section.

Molar Heat of Freezing and Condensation

You may have wondered why table 10.3.1 only provides the molar heat of fusion and vaporization, and not the opposite processes of freezing and condensation. For example, if we wanted to quantify the amount of heat it takes to freeze a certain mass of N_2, what would be the molar heat of freezing? This is actually the negative of molar heat of melting or fusion, which is the heat energy required to melt one mole of a substance. So from the above diagram, it is - 0.71 kJ/mol. Note, this is a consequence of the First Law of Thermodynamics, the law of the conservation of energy, which we will discuss in detail in the next section. What it is stating is that the energy you must add to melt a mole of a substance is the same as the energy released when a mole freezes, because energy is conserved, but the energy flows in the opposite direction. Therefore, the above table not only gives you the enthalpies of fusion and vaporization, but also of freezing and condensation.

ΔH_{Fusion} = - $\Delta H_{freezing}$
$\Delta H_{vaporization}$ = -$\Delta H_{condensation}$

Molar Heat of Sublimation and Condensation

Table 10.3.1 also provides the heats of sublimation and condensation. Can you Figure them out?

$\Delta H_{Sublimation}$ = ΔH_{Fusion} + $\Delta H_{Vaporization}$
$\Delta H_{Deposition}$ = - $\Delta H_{Vaporization}$ -ΔH_{Fusion}

Note: Quantifying Heat Transfer within a Phase

The heat being pumped into the system does not always go toward changing the phase of a substance. It may go into changing the temperature of a substance. For example, if you want to melt an ice cube that is at -10°C, you have to heat the ice to 0°C, before you can actually start the melting process. In order to quantity the amount of heat required to change the temperature (in the same phase, ice) from -10°C to 0°C, you can use the specific heat capacity equation that we covered in 5.2. (specific heat capacity, q = mCΔT (section 5.2) . Remember, the specific heat capacity (c), is different depending on which phase that the substance is in).

Heating Curves

A heating curve is a graph of the temperature of a substance as a function of the amount of heat added (moving to the right) or removed (moving to the left), as depicted in Figure 5.3.2. Note the flat regions represent phase changes, and the slope of the line in the rising regions is related to the capacitance of the sample being measured (it is actually the reciprocal, or 1/mc).

The concept of the heating curve is very important as it helps us Figure out what equation to use to quantify the heat associated with phase changes and temperature changes (watch the two videos below). At the beginning of this section, we stated that two of the three things that could happen when you add heat to a substance is that either it will undergo a phase change, or its temperature will rise. These are the two effects being described by the heating/cooling curve. The mathematical equations are:

q=mc ΔT (energy of a temperature change within a phase)
q=n $\Delta H_{transition}$ (energy of a phase transition)

It needs to be realized that if you add heat, you move to the right, and if you remove heat, you move to the left.

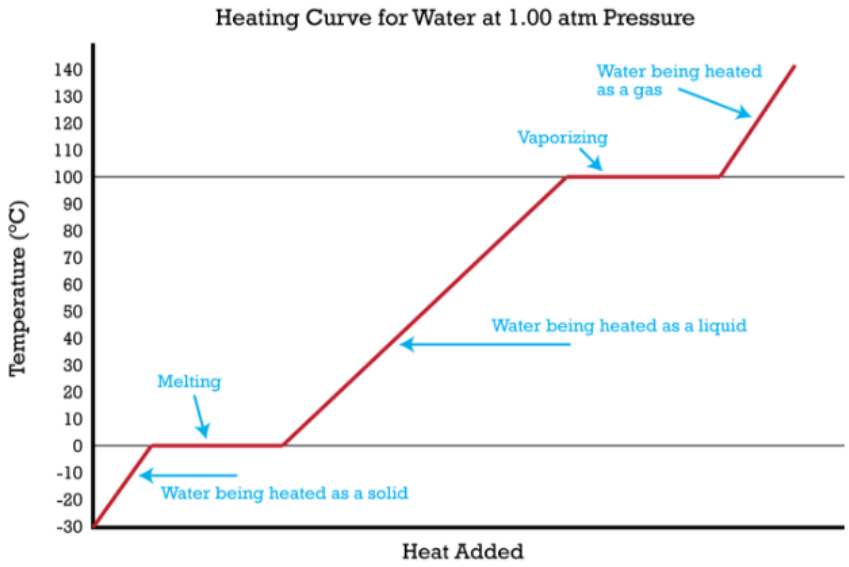


Figure 10.3.2: Cooling Curve for Water. Note, Increasing heat moves to the left, removing heat moves to the right.

What is wrong with the slopes of the lines in Figure 10.3.2?

For water c(liq) = 4.184(J/g-degC), c(ice) = 2.06(J/g-degC) and c(steam)=2.01(J/g-degC).

The slopes would be different, and the slopes of the ice and steam would roughly be twice as steep as the water. That is, it takes more energy to raise the temperature of liquid water by 1 deg than it does for either steam or ice.

What is wrong with the flat region of Figure 10.3.2?

They are roughly equal in length. In exercise 10.3.2 it was pointed out that $\Delta H_{Vaporization} >> \Delta H_{Fusion}$ and so the flat region for vaporizing should be much longer than the one for melting. As this is such an effect, a zig-zag is often drawn into the flat region to indicate that it is off the scale of the graph (see videos below).

Combined Temperate and Phase Change Problems

The following three examples deal with adding different amounts of energy to 100.0g of ice starting at -10.0°C.

Example 10.3.1 Calculation of Final Temperature

What is the final temperature when **1.6 kJ of heat** is added to 100.0 grams of ice at -10.0°C?

Solution

This problem requires you figuring out what phase the substance ends up in. Does 1.6kJ have enough enough energy to bring it to the freezing point? To melt it? To warm the liquid to its boiling point? To boil it? To raise the energy of a super heated steam?

YouTube Video

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Video 10.3.3: 5'35" YouTube showing the final temperature after adding 1.6kJ of heat to 100.0g ice at -10.0°C, (https://youtu.be/pjhZVhFjHGc).

Example 10.3.2 Calculation of Final Temperature

What is the final temperature when **25.00 kJ of heat** is added to 100.0 grams of ice at -10.0°C?

Solution

This problem requires you figuring out what phase the substance ends up in. Does 25.00kJ have enough enough energy to bring it to the freezing point? To melt it? To warm the liquid to its boiling point? To boil it? To raise the energy of a super heated steam?

YouTube Video

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Video 10.3.3: 3'08" YouTube showing the final temperature after adding 25.00kJ of heat to 100.0g ice at -10.0°C, (https://youtu.be/3cjSloZOsHQ).

Exercise 10.3.3

After all the ice has melted in the above problem, is there more ice, or liquid water?

Answer available at the end of this section.

Example 10.3.3 Calculation of Final Temperature

What is the final temperature when **310.0 kJ of heat** is added to 100.0 grams of ice at -10.0°C?

Solution

This problem requires you figuring out what phase the substance ends up in. Does 310.0kJ have enough enough energy to bring it to the freezing point? To melt it? To warm the liquid to its boiling point? To boil it? To raise the energy of a super heated steam?

After adding 310 kJ you have enough energy to raise the super cooled water from -10.0^{o}C to its freezing point (2.06kJ), melt the ice (33.42kJ), heat the liquid water to its boiling point (41.84kJ) and boil all the water (225.37kJ) using a total of 303.7 kJ, with 7.3kJ left to raise the temperature of the super-heated steam to a final temperature of 136^{o}C. Notice how the majority of the energy was used in the fourth step, the boiling of the the water (225.37kJ out of 310kJ, or 72.7% of the entire energy). This makes sense, as in going from a liquid to a gas you are overcoming all of the forces holding the water molecules together.

YouTube Video

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Video 10.3.3: 3'08" YouTube showing the final temperature after adding 310.0 kJ of heat to 100.0g ice at -10.0°C, (https://youtu.be/3cjSloZOsHQ).

Test Yourself

Homework: Section 10.3

Query 10.3.1

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Query 10.3.2

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- Ronia Kattoum (UALR)
- Liliane Poirot

Answers

Exothermic. It is easier for us to think about the opposite process of melting. It takes energy to melt ice, so heat is flowing into the system(the ice cube) and out of the surroundings (the air, let's say). So the surroundings (air) during the process of melting the ice. The opposite of melting ice is freezing water. So, you expect this process to be exothermic, releasing heat to the surroundings, and therefore, warming the surrounds.

Warm

From table 10.3.1 it is clear that the enthalpy of vaporization is larger than fusion. This is because when you vaporize something you overcome all of the intermolecular forces and the material becomes a gas, filling up the entire container. It takes much more energy to vaporize a substance then to melt it.

Since it takes 33.24kJ to melt all the ice, and you have 22.94 left over after bring the supercooled ice to its freezing point, you have effectively applied the 22.94

kJ to the ice at zero degrees, and so the fraction melted is what you applied over what was required to melt it all. So 22.94/33.24 = 0.69, so it is 69% liquid water and 31% ice (69 grams of the 100 grams melted, and 31 are still frozen).

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10.4: First Law of Thermodynamics

Learning Objectives

- Determine how heat and work are related to the overall internal energy of a system
- Calculate the work done by/on a system at constant pressure (PV work)
- Calculate changes in internal energy and enthalpy using the First Law of Thermodynamics
- Define a state function and determine which variables are state functions

Introduction

In the very first Chapter of this book we described three attributes of matter.

- Matter has Mass
- Matter has Volume
- Matter has Energy

Thermodynamics comes from the Greek words *therme* (heat) and *dynamis* (power) and deals with the third attribute of matter, energy. In section 5.1.2 we defined energy as the capacity to do work or transfer heat. The **first law of thermodynamics** states that energy can not be created or destroyed, that is, the energy of the universe is constant. This allows us to define a thermodynamic system, which in the case of chemistry is a reaction, with the surroundings being everything else.

So what do we mean the reaction is a thermodynamic system? Lets consider the combustion of hydrogen

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O \tag{10.4.1}$$

During a chemical reaction the energy of the system changes as bonds in reactants are broken and the atoms rearrange to form the new bonds of the products. These new bonds are typically higher or lower in energy. If the products store more energy in their bonds, that energy must come from somewhere, and so it is transferred into the system from the surroundings. In the case of combustion, the products are lower in energy and so that energy is transferred to the surroundings. The key point here is as a reaction proceeds energy flows into or out of the system, depending on whether energy is absorbed or released as the products are formed. So what is positive energy flow?

Convention for Energy Flow

The convention thermodynamics use is that positive energy is when the system gains energy, which is when either heat is transferred to, or work is done on the system. This requires differentiating the system from the surroundings. In chemical thermodynamics, the system represents the internal energy of the atoms of the reactants as they become products.

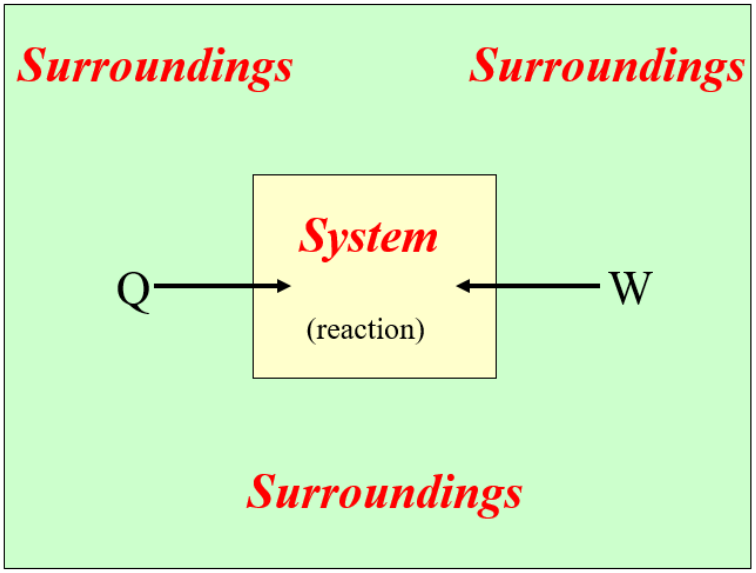


Figure 10.4.1: Diagram showing the convention for the flow of positive energy.

- Positive Energy Flow is when heat is added and/or work done to the system. (the system gains energy)
  - $\Delta E_{reaction} > 0$
- Negative Energy Flow is when the system releases work to the surroundings and/or does work on the surroundings. (the system looses energy).
  - $\Delta E_{reaction} < 0$

From the perspective of a thermodynamic system we can view the universe to consist of the system and the surroundings

$$Universe = System + Surroundings \tag{10.4.2}$$

The first law states that the energy of the universe is a constant and so the change in the energy of the universe must equal zero.

$$\Delta E_{Universe} = 0 \tag{10.4.3}$$

Since the Universe is the system and the surrounding this means the change in energy of the system and surrounding must also equal zero

$$\Delta E_{Universe} = \Delta E_{System} + \Delta E_{Surrounding} = 0 \tag{10.4.4}$$

therefore,

$$\Delta E_{System} = -\Delta E_{Surrounding} \tag{10.4.5}$$

This is logical and means any energy lost by the system must be gained by the surroundings, and vice-versa, any energy gained by the system must be lost by the surroundings.

We define the energy of the system as the internal energy and give it the symbol U.

$$\Delta U = -\Delta E_{Surrounding} \tag{10.4.6}$$

Note: Direction of Heat Transfer

Students often confuse heat and temperature. If two objects are in thermal contact with each other and are different temperatures, heat will transfer from the higher temperature object to the one at the lower. So the temperature difference defines the direction of heat transfer, but not the energy of the ojbects. To understand this we need to look at the Second Law of Thermodynamics and another state function called entropy, which is covered in the second semester class.

Exercise 10.4.1

Does heat always flow from the object with more energy to the one with less?

Answer available at the end of this section.

Calorimetry

Calorimetery is an application of the First Law of Thermodynamics to heat transfer, and allows us to measure the enthalpies of reaction or the heat capacities of substances. From the first law we can state

$$\Delta E_{Universe} = \Delta E_{System} + \Delta E_{Surrounding} = 0 \tag{10.4.7}$$

therefore,

$$\Delta E_{System} = -\Delta E_{Surrounding} \tag{10.4.8}$$

A calorimeter is an instrument that has a thermally isolated compartment that is designed to prevent heat flow to the surroundings. Technically, this is called an adiabatic surface in that no heat flows in and out of the calorimeter, which is in contrast to isothermal, which means constant temperature. In truth, there is no such thing as a thermally isolated system, but we can design systems that approach it on the time scale of a measurement. We will consider two types of calorimeters, ideal calorimeters, where the calorimeter does not absorb heat itself, and real calorimeters, where the calorimeter absorbs heat. We will apply calorimetry to two different thermodynamic processes, that of transference of heat across objects at different temperatures, and that released or absorbed by a chemical reaction. We will look at two ways of measuring this heat, either the temperature change of the calorimeter, or the mass of a substance that undergoes a phase change within the calorimeter at a constant temperature.

If we consider the work in a calorimeter to be zero, and noting that Δ E = Δ Q + W, the above equation becomes

$$\Delta Q_{system} = -\Delta Q_{Surrounding} \tag{10.4.9}$$

If we can measure the heat absorbed or lost by the surroundings we can use the above equivalence statement to calculate the heat lost or absorbed the system, and this is the basis for calorimetry. Figure 10.4.1 shows the basic layout of an ideal calorimeter where the gray shaded box acts as a surface that prevent any heat entering or leaving the solution, and so the solution functions as the surroundings. In the first exothermic process heat is released by the system and the temperature rises, while in the second endothermic process heat is absorbed by the system and the temperature decreases.

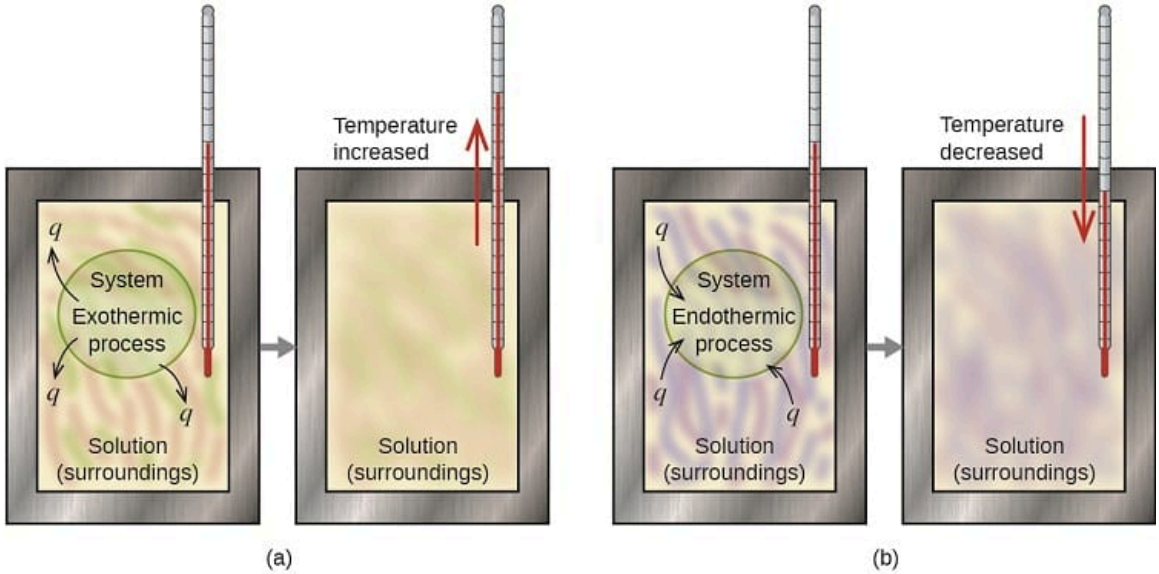


Figure 10.4.1: In part a, and exothermic process occurs and the calorimeter's temperature rises, while in part b an endothermic process occurs and the temperature falls.

Heat Transfer Between Hot and Cold Objects in an Ideal Calorimeter

When a hot object (T_H) is placed in thermal contact with a cold object (T_C) energy is transferred between the two, but the rate of heat transfer from the hot to the cold is greater, so it cools down while the cold object heats up, until they reach the same temperature (T_F). Once they are at the same temperature the system is at thermal equilibrium and the rate of transfer of energy from the hot to cold is the same as cold to hot. That is, energy is a vectoral flow, and even at the same temperature, there is a flow of energy across the boundary of two objects, but the flow is the same.

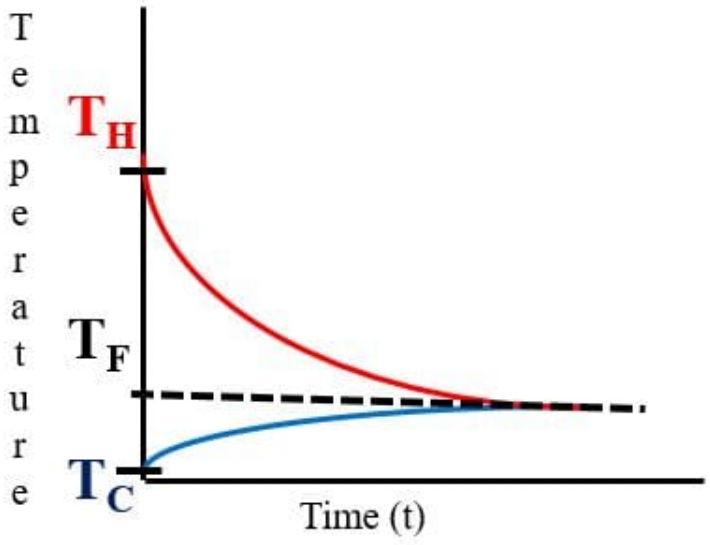


Figure 10.4.2: Graph showing temperature change as a function of time after a hot object is put in thermal contact with a cold object. Note, that in this case the capacity of the cold object to hold heat is greater, so the change in temperature of the cold object is less than the hot. Also note, $\Delta T = T_{Final} - T_{Initial}$ where $T_{Initial} = T_{Hot}$ for the hot object and $T_{Initial} = T_{Cold}$ for the cold object, so ΔT_{Hot} is a negative number (it cools, so $T_F < T_i$), while ΔT_{Cold} is a positive number (it heats, so $T_F > T_i$).

Starting with the first law for two isolated objects and considering no work, we have

$$q_{cold} + q_{hot} = 0 \tag{10.4.10}$$

so

$$q_{cold} = -q_{hot} = 0 \tag{10.4.11}$$

> The amount of heat lost by a warmer object equals the amount of heat gained by a cooler object.

From section 5.2 we know that for a substance in a constant phase the heat relates to the temperature by equation 5.2.3 (q=mc\(\Delta|)T, which substituting into the above two equations give

$$m_C c_C \Delta T_C + m_H c_H \Delta T_H = 0 \tag{10.4.12}$$

or

$$m_C c_C \Delta T_C = -m_H c_H \Delta T_H \tag{10.4.13}$$

where

$$\Delta T_H = T_F - T_H \;\; and \;\; \Delta T_C = T_F - T_C \tag{10.4.14}$$

In an ideal calorimeter we consider the calorimeter to absorb no heat itself, so if we have material of a uniform composition, we can relate the heat transferred from the hot object to the cold object by eq. 5.6.7.

Example 10.4.1 Ideal calorimeter problem

What is T_F when 2.00g Au at 100.0°C is placed into 25.0 ml H_2O at 30.0°C? You can obtain the specific heat capacities for gold and water from table 5.2.1.

Solution

You are solving 5.6.7 for T_F.

$$T_F = \frac{m_H c_H T_H + m_c c_c T_c}{m_H c_H + m_c c_c} = 30.17^o C \tag{10.4.15}$$

The solution to this problem is worked out in video 10.4.1

YouTube Video

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View this YouTube Video online

Video 10.4.1: 7'05" YouTube solving example 10.4.1 (https://youtu.be/Htqzz9TAfrQ)

Enthalpy

Enthalpy (H) comes from the Greek *enthalpein*, meaning ("to warm"). which is the measure of the heat of a system. From above we know the change in internal energy (U) of a system is equal to the heat and work done on the system.

$$\Delta U = Q + W \tag{10.4.16}$$

From section 5.1 we know the work of expansion is PV work, which for a system at constant pressure is

$$W = -P\Delta V \tag{10.4.17}$$

Enthalpy ΔH is defined as the heat transferred at constant pressure Q_p,

$$\Delta H = Q_p \tag{10.4.18}$$

substituting the definitions of work (Equation 10.4.17) and heat (Equation 10.4.18) into Equation 10.4.16 at constant P gives,

$$\Delta U = \Delta H - P\Delta V \tag{10.4.19}$$

rearranging to solve for enthalpy gives:

$$\Delta H = \Delta U + P\Delta V \tag{10.4.20}$$

a more general definition of enthalpy (which can be extended to conditions that are not at constant pressure is

$$H = U + PV \tag{10.4.21}$$

Note: relating enthalpy to work of expansion

In section 5.1.4 we introduced PV work as the work of expansion. That is, an expanding gas does work against the environment, and it takes work to compress a gas. According to the First Law, Positive work is when work is done on the system, which is that of compression, where Δ V is negative. so -PΔV is positive work, and equation 10.4.20 becomes

$$\Delta H = \Delta U - W \tag{10.4.22}$$

That is the enthalpy of a reaction is the heat transferred minus the work done on the system by a change in the volume of the system. That is, some of the energy change in going from reactants to products can be consumed by the system as it changes volumes, and this is most noticeable in gas phase systems.

Exercise 10.4.2

Calculate the Enthalpy change for a gas phase reaction at 1.2 atm if the change in internal energy is 47kJ and the volume increase by 1.1 L to 41.0 L. Note 1L-atm =101.325J

Answer available at the end of this section.

We often call enthalpy the Heat of Reaction and tabulate tables of enthalpies of reaction, which can be used to tabulate the heat given off during a reaction.

State Functions

Enthalpy and Internal Energy are state functions. A state function is a function whose value is defined by the state, its pressure, temperature, composition, and amount of substance, and not the process of how it was obtained. So the difference between two state functions is path independent. Thus any solution of NaCl at 25°C and 1 atm that contains a mixture of 1 mol NaCl and 100 mol H_2O has the same internal energy and the same enthalpy as any other solution with the same specifications. It does not matter whether the solution was prepared by simply dissolving NaCl(s) in H_2O, by reacting NaOH(*aq*) with HCl(*aq*), or by some other method.

Is work a state function? No, as work is path dependent. If you look at Figure 10.4.2 you have a nice analogy for understanding the state function. The gravitational potential energy difference between an object at a height of the summit of Mt. Kilimanjaro and it's base is the same, no matter how you made it to the summit. The green and red path "X" and "Y" represent different paths, and the work required does depend on the path.

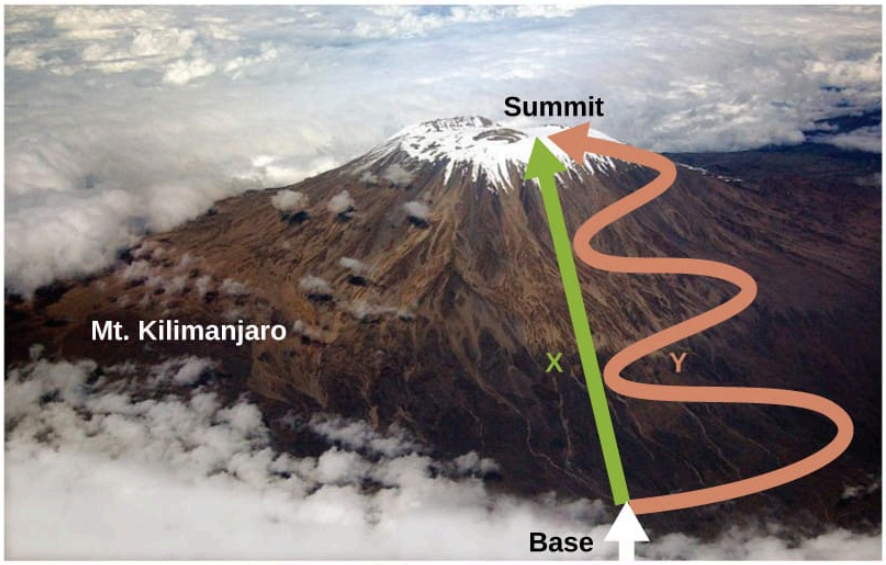


Figure 10.4.2: Diagram showing difference between path dependent function and a state function. The work required to reach the summit is not a state function, and depends on the path, but once you are on the top, the view is the same, no matter how you got there.

Test Yourself

Homework: Section 10.4

Query 10.4.1

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- Ronia Kattoum (UALR)
- anonymous
- Liliane Poirot

Answers

No, it transfers from the one with the higher temperature to the one at a lower temperature. A lake at 30°C has more energy than a cup of water at 35°C, but if you add the water to the lake, it will cool as energy flows from the water to the lake.

$$\Delta H = 47kJ + 1.2atm(41.0L - 1.1L)\left(\frac{101.325J}{L \cdot atm}\right) = 52kJ \tag{10.4.23}$$

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10.5: Enthalpy of Reaction

Learning Objectives

- Recognize that breaking bonds requires energy, and making new bonds releases energy
- Assess if the overall bond making and breaking process results in net energy input or output leading to overall exothermic or endothermic reactions
- Define standard reaction enthalpy and relate to stoichiometric coefficients in a a balanced chemical reaction
- Calculate the heat produced/consumed in chemical reaction based on amount of a reactants and products using standard reaction enthalpy and vice versa

Enthalpy of Reaction

When a reaction occurs, bonds in reactants break, the atoms rearrange and new bonds form as the products are created. If the bond energies of the formed bonds are different than the bond energies of the broken bonds, then the energy of the reactants is not equal to the energy of the products, and energy is either absorbed or released into the surroundings. There are a few things to note.

- Breaking bonds always requires energy, and so it is endothermic.
- Forming bonds is the opposite of breaking, and so it is exothermic.

Because enthalpy is a state function, we can energetically view a reaction as taking place in two steps (as illustrated in Figure 10.5.1. The first step is the endothermic step of breaking the reactant bonds and the second is the exothermic step of making the product bonds. The energy change of the reaction can be viewed as the sum of these two steps, and results in two possibilities,

1. The exothermic step (2) is greater than the endothermic step (1) and the reaction is exothermic.
2. The endothermic step (1) is greater than the exothermic step (2) and the reaction is endothermic.

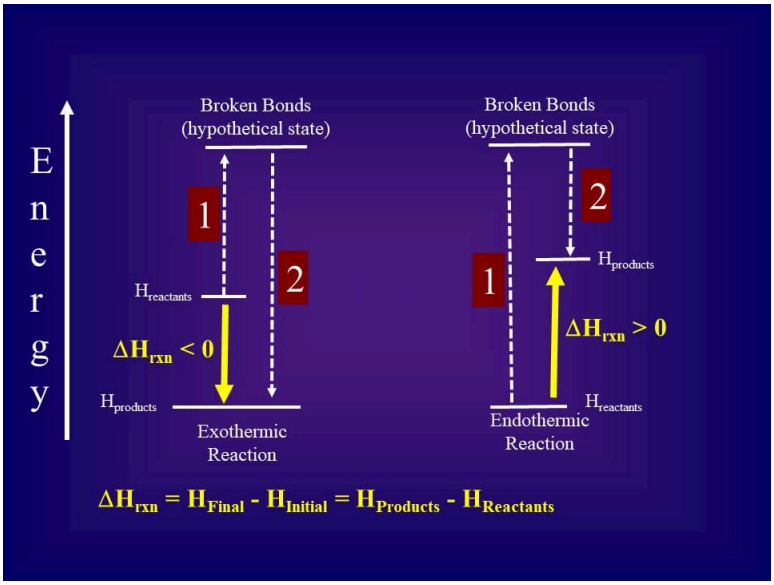


Figure 10.5.1: Energy cycle diagrams showing exothermic and endothermic reactions in terms of an initial endothermic step involving breaking reactant bonds, followed by an exothermic step involving making product bonds.

Note, we are interested in ΔH(reaction), which only depends on the initial (reactant) and final (product) state, because enthalpy is a state function this difference is path independent. What we do know is that the first step required energy and the second released energy. In Figure 10.5.2, we are only focusing on the energy difference between the initial and final states (yellow transition of Figure 10.5.1).

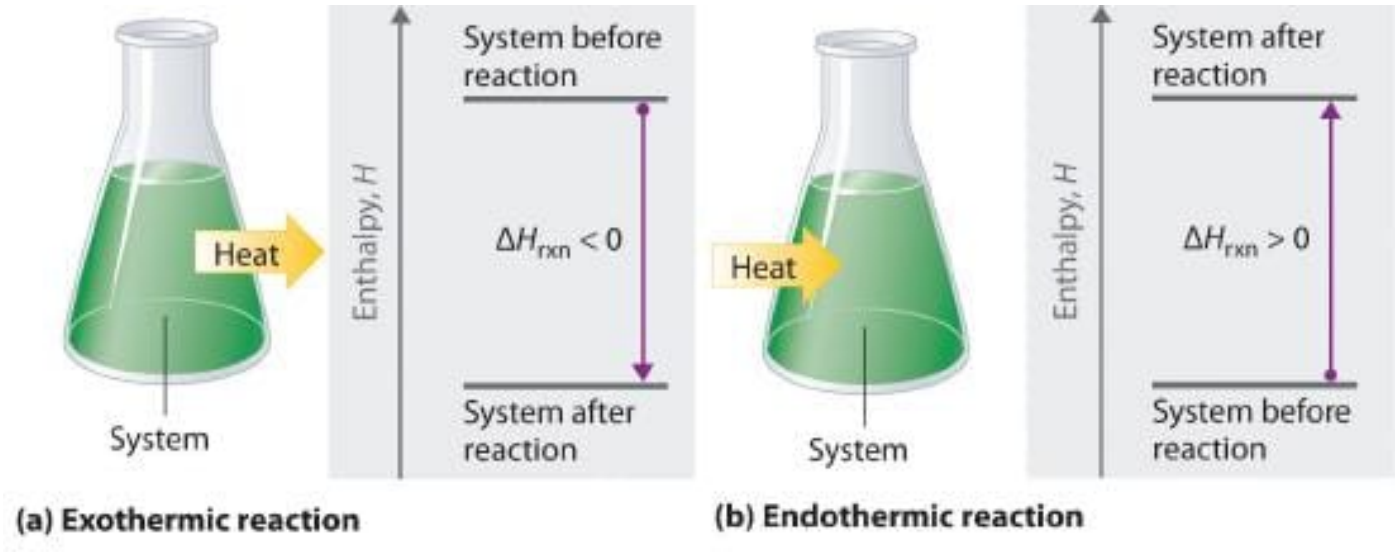


Figure 10.5.2: In this image we are only focusing on the reactant and product states, not the path. On the left side, the energy described the the enthalpy of reaction is released to the environment, and it warms the environment. On the right side, the system absorbs energy, cooling the environment.

Enthalpies of Reaction

In Chapter 3 (section 3.1) we learned about Lavoisier's "Law of Conservation of Matter", and from this we learned that chemical equation must be balanced. In the next Chapter 4, we applied stoichiometry to balanced equations and were able to predict reaction yields based on the complete consumption of the limiting reagent. In this section we are going to apply the First Law of Thermodynamics, the conservation of energy, to the balanced equation, and relate the energy of a chemical reaction to the stoichiometric coefficients of a reaction. The Enthalpy of Reaction is the energy evolved when the number of moles of each species is their coefficient, and this results in an equivalence statement that relates the stoichiometric number of each species to the enthalpy of reaction. The enthalpy of reaction is not written as per mole, but that information is implicit.

Enthalpy and Chemical Reactions

Endothermic Reactions: In endothermic reactions the heat is added to the system as reactants convert to products, and so the sign of ΔH is positive. Now the units of $\Delta H_{reaction}$ often confuse students, as there are actually implicit units of mole in the denominator, where it is based on the stoichiometric coefficient of the chemical you are interested in. The easiest way to look at this is as an equivalence statement, where the enthalpy of reaction is related to the coefficient of each species.

Reactants + Heat → Products, ΔH>0

Example:

$$4CO_2(g) + 2H_2O(l) \rightarrow 2C_2H_2(g) + 5O_2(g) \quad \Delta H_{reaction} = 2599kJ \tag{10.5.1}$$

The above equation is really saying 2599 kJ are being released for every 4 moles of CO_2, 2 moles H_2O, 2 moles C_2H_2 and 5 moles O_2.

This gives four equivalence statements

- 4 mole CO_2 = 2559kJ
- 2 mole H_2O = 2559kJ
- 2 mole C_2H_2 = 2559kJ
- 5 mole O_2 = 2559kJ

Example 10.5.1

Calculate the energy released when 88.0g C_2H_2 react with excess oxygen according to the above equation

Solution

Note: C_2H_2 is a product and in so the above equation needs to be written backwards, which changes the sign from endothermic (+) to exothermic (-).

$2C_2H_2(g)+5O_2(g) \rightarrow 4CO_2(g)+2H_2O(l)$ ($\Delta H_{reaction}$ = -2599 kJ)

$$88.0gC_2H_2\left(\frac{molC_2H_2}{26.0378g}\right)\left(\frac{-2559kJ}{2molC_2H_2}\right) = -4,324kJ \tag{10.5.2}$$

Exothermic reactions: In exothermic reactions heat is released from the system as reactants convert to products, and so the sign of ΔH is negative. Note, in example 10.5.1 1 we wrote the endothermic reaction of 10.5.1 backwards and it became exothermic. This is an important principle and a consequence of the First Law of Thermodynamics. That is, if a reaction is endothermic in one direction, it is exothermic in the other, and the amount of energy absorbed in the endothermic direction is equal to the amount released in the exothermic direction.

Reactants + Heat → Products + Heat, ΔH<0

Example:

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l) \quad \Delta H_{reaction} = -2599kJ \tag{10.5.3}$$

Looking at the energetics as a series of equivalence statements gives:

4 mole CO_2 = -2559kJ
2 mole H_2O = -2559kJ
2 mole C_2H_2 = -2559kJ
5 mole O_2 = -2559kJ

Test Yourself

Homework: Section 10.5

Exercise 10.5.1

According to Equation 10.5.3, how much heat is evolved if 88.0 g of oxygen reacts with excess C_2H_2?

Answer available at the end of this section.

Exercise 10.5.2

Consider the following reaction:

$$\mathrm{CO + 2\,H_2 \longrightarrow CH_3OH + 128\,kJ}$$

a. Is the reaction endothermic or exothermic?
b. If 12.2 g of hydrogen gas are reacted _____kJ are _____(produced/consumed).
c. If 455 kJ of energy is produced, what was the mass of CO that was reacted?

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Query 10.5.1

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Contributors and Attributions

- Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Ronia Kattoum (UALR)
- anonymous
- Liliane Poirot

Answers

$$88.0gO_2\left(\frac{molO_2}{32.00g}\right)\left(\frac{-2559kJ}{5molO_2}\right) = -1,410kJ \tag{10.5.4}$$

Exothermic. Since the heat is written on the right side of the equation, it can be thought of as heat produced or generated. So, it is an exothermic reaction. This reaction can also be written in this format, where the enthalpy is represented as having a negative value.

387 kJ are produced:

99.6 grams of CO reacted:

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10.6: Enthalpy Calculations

Learning Objectives

- Define Hess's Law and relate it to the first law of thermodynamics and state functions
- Calculate the unknown enthalpy of a reaction from a set of known enthalpies of combustion using Hess's Law
- Define molar enthalpy of formation of compounds
- Calculate the molar enthalpy of formation from combustion data using Hess's Law
- Using the enthalpy of formation, calculate the unknown enthalpy of the overall reaction
- Calculate the heat evolved/absorbed given the masses (or volumes) of reactants.

Hess's Law

Hess's Law states that if you can add two chemical equations and come up with a third equation, the enthalpy of reaction for the third equation is the sum of the first two. This is a consequence of the First Law of Thermodynamics, the fact that enthalpy is a state function, and brings for the concept of coupled equations.

Coupled Equations: A balanced chemical equation usually does not describe how a reaction occurs, that is, its mechanism, but simply the number of reactants in products that are required for mass to be conserved. In reality, a chemical equation can occur in many steps with the products of an earlier step being consumed in a later step. For example, consider the following reaction phosphorous reacts with oxygen to from diphosphorous pentoxide ($2P_2O_5$)

$$P_4 + 5O_2 \rightarrow 2P_2O_5 \tag{10.6.1}$$

and then the product of that reaction in turn reacts with water to form phosphorus acid

$$2P_2O_5 + 6H_2O \rightarrow 4H_3PO_4 \tag{10.6.2}$$

In the above equation the P_2O_5 is an intermediate, and if we add the two equations the intermediate can cancel out. Hess's law states that if two reactions can be added into a third, the energy of the third is the sum of the energy of the reactions that were combined to create the third.

$$\begin{array}{lll} \text{equation 1:} & P_4 + 5O_2 \rightarrow 2P_2O_5 & \Delta H_1 \\ \text{equation 2:} & 2P_2O_5 + 6H_2O \rightarrow 4H_3PO_4 & \Delta H_2 \\ \\ \text{equation 3:} & P_4 + 5O_2 + 6H_2O \rightarrow 3H_3PO_4 & \Delta H_3 \end{array} \tag{10.6.3}$$

Enthalpy is a **state function** which means the energy change between two states is independent of the path. Since equation 1 and 2 add to become equation 3, we can say:

$$\Delta H_3 = \Delta H_1 + \Delta H_2 \tag{10.6.4}$$

Hess's Law says that if equations can be combined to form another equation, the enthalpy of reaction of the resulting equation is the sum of the enthalpies of all the equations that combined to produce it.

Stepwise Calculation of ΔH_f°

Using Hess's Law Determine the enthalpy of formation, ΔH_f°, of $FeCl_3(s)$ from the enthalpy changes of the following two-step process that occurs under standard state conditions:

$$\mathrm{Fe}(s) + \mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_2}(s) \quad \Delta H^\circ = -341.8\ \mathrm{kJ}$$

$$\mathrm{FeCl_2}(s) + \frac{1}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) \quad \Delta H° = -57.7\ kJ$$

Solution

We are trying to find the standard enthalpy of formation of FeCl3(s), which is equal to ΔH° for the reaction:

$$\mathrm{Fe}(s) + \frac{3}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) \quad \Delta H_\mathrm{f}^\circ = ?$$

Looking at the reactions, we see that the reaction for which we want to find Δ*H*° is the sum of the two reactions with known Δ*H* values, so we must sum their Δ*H*s:

$$\begin{array}{ll} \mathrm{Fe}(s) + \mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_2}(s) & \Delta H° = -341.8\ \mathrm{kJ} \\ \mathrm{FeCl_2}(s) + \frac{1}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) & \Delta H° = -57.7\ \mathrm{kJ} \\ \hline \mathrm{Fe}(s) + \frac{1}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) & \Delta H° = -399.5\ \mathrm{kJ} \end{array}$$

The enthalpy of formation, $\Delta H_\mathrm{f}^\circ$, of $FeCl_3$(s) is –399.5 kJ/mol.

Exercise 10.6.1

Calculate ΔH for the process:

$$\mathrm{N_2}(g) + 2\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{NO_2}(g)$$

from the following information:

$$\mathrm{N_2}(g) + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{NO}(g) \quad \Delta H = 180.5\ \mathrm{kJ}$$

$$\mathrm{NO}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{NO_2}(g) \quad \Delta H = -57.06\ \mathrm{kJ}$$

Answer available at the end of this section.

Here is a less straightforward example that illustrates the thought process involved in solving many Hess's law problems. It shows how we can find many standard enthalpies of formation (and other values of Δ*H*) if they are difficult to determine experimentally.

A More Challenging Problem

Using Hess's Law Chlorine monofluoride can react with fluorine to form chlorine trifluoride:

$$\text{(i)}\ \mathrm{ClF}(g) + \mathrm{F_2}(g) \longrightarrow \mathrm{ClF_3}(g) \quad \Delta H\,° = \ ?$$

Use the reactions here to determine the ΔH° for reaction (i):

$$\text{(ii)}\ 2\,\mathrm{OF_2}(g) \longrightarrow \mathrm{O_2}(g) + 2\,\mathrm{F_2}(g) \quad \Delta H_{(ii)}^\circ = -49.4\ \mathrm{kJ}$$

$$\text{(iii)}\ 2\,\mathrm{ClF}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{Cl_2O}(g) + \mathrm{OF_2}(g) \quad \Delta H_{(iii)}^\circ = +205.6\ \mathrm{kJ}$$

$$\text{(iv)}\ \mathrm{ClF_3}(g) + \mathrm{O_2}(g) \longrightarrow \tfrac{1}{2}\mathrm{Cl_2O}(g) + \frac{3}{2}\mathrm{OF_2}(g) \quad \Delta H_{(iv)}^\circ = +266.7\ \mathrm{kJ}$$

Solution

Our goal is to manipulate and combine reactions (ii), (iii), and (iv) such that they add up to reaction (i). Going from left to right in (i), we first see that $\mathrm{ClF}_{(g)}$ is needed as a reactant. This can be obtained by multiplying reaction (iii) by $\frac{1}{2}$, which means that the ΔH° change is also multiplied by $\frac{1}{2}$:

$$\mathrm{ClF}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \frac{1}{2}\mathrm{Cl_2O}(g) + \frac{1}{2}\mathrm{OF_2}(g) \quad \Delta H° = \frac{1}{2}(205.6) = +102.8\ k\,\mathrm{J}$$

Next, we see that $\mathrm{F_2}$ is also needed as a reactant. To get this, reverse and halve reaction (ii), which means that the ΔH° changes sign and is halved:

$$\frac{1}{2}\mathrm{O_2}(g) + \mathrm{F_2}(g) \longrightarrow \mathrm{OF_2}(g) \quad \Delta H° = +24.7\ k\,\mathrm{J}$$

To get $\mathrm{ClF_3}$ as a product, reverse (iv), changing the sign of ΔH°:

$$\frac{1}{2}\mathrm{Cl_2O}(g) + \frac{3}{2}\mathrm{OF_2}(g) \longrightarrow \mathrm{ClF_3}(g) + \mathrm{O_2}(g) \quad \Delta H° = -266.7\ \mathrm{kJ}$$

Now check to make sure that these reactions add up to the reaction we want:

$$\begin{array}{ll}
\mathrm{ClF}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \frac{1}{2}\mathrm{Cl_2O}(g) + \frac{1}{2}\mathrm{OF_2}(g) & \Delta H° = +102.8\ \mathrm{kJ} \\
\frac{1}{2}\mathrm{O_2}(g) + \mathrm{F_2}(g) \longrightarrow \mathrm{OF_2}(g) & \Delta H° = +24.7\ \mathrm{kJ} \\
\frac{1}{2}\mathrm{Cl_2O}(g) + \frac{3}{2}\mathrm{OF_2}(g) \longrightarrow \mathrm{ClF_3}(g) + \mathrm{O_2}(g) & \Delta H° = -266.7\ \mathrm{kJ} \\
\hline
\mathrm{ClF}(g) + \mathrm{F_2} \longrightarrow \mathrm{ClF_3}(g) & \Delta H° = -139.2\ \mathrm{kJ}
\end{array}$$

Reactants $\frac{1}{2}\mathrm{O_2}$ and $\frac{1}{2}\mathrm{O_2}$ cancel out product O_2; product $\frac{1}{2}\mathrm{Cl_2O}$ cancels reactant $\frac{1}{2}\mathrm{Cl_2O}$; and reactant $\frac{3}{2}\mathrm{OF_2}$ is cancelled by products $\frac{1}{2}\mathrm{OF_2}$ and OF_2. This leaves only reactants ClF(g) and F_2(g) and product ClF_3(g), which are what we want. Since summing these three modified reactions yields the reaction of interest, summing the three modified ΔH° values will give the desired ΔH°:

$$\Delta H^\circ = (+102.8\ \mathrm{kJ}) + (24.7\ \mathrm{kJ}) + (-266.7\ \mathrm{kJ}) = -139.2\ \mathrm{kJ}$$

Exercise 10.6.2

Aluminum chloride can be formed from its elements:

$$\text{(i)}\ 2\,\mathrm{Al}(s) + 3\,\mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{AlCl_3}(s) \quad \Delta H^\circ = \ ?$$

Use the reactions here to determine the ΔH° for reaction (i):

$$\text{(ii)}\ \mathrm{HCl}(g) \longrightarrow \mathrm{HCl}(aq) \quad \Delta H^\circ_{(ii)} = -74.8\ \mathrm{kJ}$$

$$\text{(iii)}\ \mathrm{H_2}(g) + \mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{HCl}(g) \quad \Delta H^\circ_{(iii)} = -185\ \mathrm{kJ}$$

$$\text{(iv)}\ \mathrm{AlCl_3}(aq) \longrightarrow \mathrm{AlCl_3}(s) \quad \Delta H^\circ_{(iv)} = +323\ \mathrm{kJ/mol}$$

$$\text{(v)}\ 2\,\mathrm{Al}(s) + 6\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{AlCl_3}(aq) + 3\,\mathrm{H_2}(g) \quad \Delta H^\circ_{(v)} = -1049\ \mathrm{kJ}$$

Answer available at the end of this section.

We also can use Hess's law to determine the enthalpy change of any reaction if the corresponding enthalpies of formation of the reactants and products are available.

Calculating Enthalpy of Reaction from Combustion Data

In section 5.6.3 we learned about bomb calorimetry and enthalpies of combustion, and table 10.6.1 contains some molar enthalpy of combustion data. In this section we will use Hess's law to use combustion data to calculate the enthalpy of reaction for a reaction we never measured. In fact, it is not even a combustion reaction.

Compound	Heat of Combustion (kJ/mol)
Hydrogen (H_2)	-286
Methane (CH_4)	-889
Ethane (C_2H_6)	-1,560
Propane (C_3H_8)	-2,220
Carbon	-393.5
Ethylene (C_2H_4)	-1411

Table 10.6.1 Heats of combustion for some common substances. Note, these are negative because combustion is an exothermic reaction. They are often tabulated as positive, and it is assumed you know they are exothermic. Also, these are not reaction enthalpies in the context of a chemical equation (section 5.5.2), but the energy per mol of substance combusted.

Note: tips for hess's law calculations

The following tips should make these calculations easier to perform

1. Write the equation you want on the top of your paper, and draw a line under it.
2. From data tables find equations that have all the reactants and products in them for which you have enthalpies.
3. If an equation has a chemical on the opposite side, write it backwards and change the sign of the reaction enthalpy.
4. If the equation has a different stoichiometric coefficient than the one you want, multiply everything by the number to make it what you want, including the reaction enthalpy

Watch Video 10.6.1 to see these steps put into action while solving example 10.6.1.

Example 10.6.3 Calculating enthalpy of reaction with hess's law and combustion table

Using table 10.6.1 Calculate the enthalpy of reaction for the hydrogenation of ethene into ethane

$$C_2H_4 + H_2 \rightarrow C_2H_6$$

Solution

Hess's Law is a consequence of the first law, in that energy is conserved.

We want the enthalpy for the reaction

$$C_2H_4 + H_2 \rightarrow C_2H_6$$

From table 10.6.1 we obtain the following enthalpies of combustion

$$\begin{array}{lll} \text{eq. 1:} & H_2 + 1/2O_2 \rightarrow H_2O & \Delta H_1 = -286kJ/mol \\ \text{eq. 2:} & C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O & \Delta H_2 = -1411kJ/mol \\ \text{eq. 3:} & C_2H_6 + 3/2O_2 \rightarrow 2CO_2 + 3H_2O & \Delta H_3 = -1560kJ/mol \end{array} \quad (10.6.5)$$

Video 10.6.1 shows how to tackle this problem. The trick is to add the above equations to produce the equation you want. So, identify species that only exist in one of the given equations and put them on the desired side of the equation you want to produce, following the Tips above. Your final answer should be -131kJ/mol

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 10.6.1: 7'00" YouTube applying Hess's law to calculate the enthalpy of hydrogenation of ethene from combustion data (https://youtu.be/-t20lkJTQYs)

Enthalpy as a State Function

What is important here, is that by measuring the heats of combustion scientists could acquire data that could then be used to predict the enthalpy of a reaction that they may not be able to directly measure. This is a consequence of enthalpy being a state function, and the path of the above three steps has the same energy change as the path for the direct hydrogenation of ethylene. We can look at this in an Energy Cycle Diagram (Figure 10.6.2).

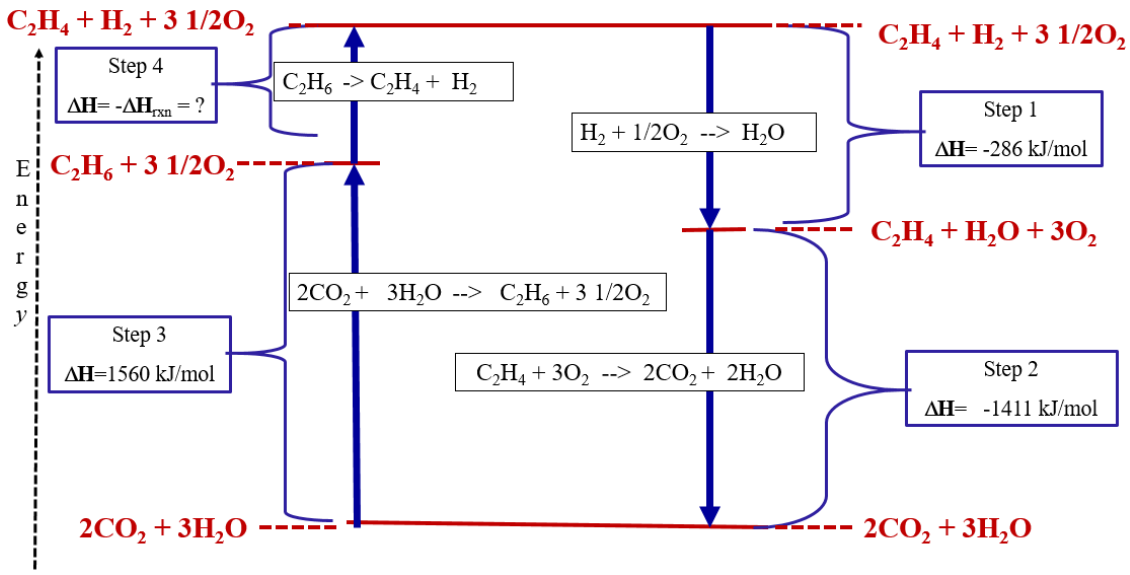


Figure 10.6.2: The steps of example 10.6.1 expressed as an energy cycle.

Because enthalpy is a state function, a process that involves a complete cycle where chemicals undergo reactions and are then reformed back into themselves, must have no change in enthalpy, meaning the endothermic steps must balance the exothermic steps.

$$\Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 = 0 \tag{10.6.6}$$

That is, the energy lost in the exothermic steps of the cycle must be regained in the endothermic steps, no matter what those steps are.

Exothermic Steps

1. ΔH_1 = -296kJ/mol
2. ΔH_2 = -1411kJ/mol Total Exothermic = -1697 kJ/mol

Endothermic Steps

3. ΔH_3 = + 1560kJ/mol

4. ΔH_4 = - ΔH^*_{rxn} = ? J/mol Total Endothermic = + 1697 kJ/mol

Note, step 4 shows C_2H_6 -- > C_2H_4 +H_2 and in example 10.6.1 we are solving for C_2H_4 +H_2 --> C_2H_6 which is the reaction of step 4 written backwards, so the answer to 10.6.1 is the negative of step 4. This problem is solved in video 10.6.1 above.

Standard Enthalpies of Formation

The **standard molar enthalpy of formation ΔH^o_f** is the enthalpy change when **1 mole of a pure substance**, **or a 1 M solute** concentration in a solution, **is formed from its elements in their most stable states under standard state conditions**. In this class, **the standard state is 1 bar and 25°C.** Note, if two tables give substantially different values, you need to check the standard states. One of the values of enthalpies of formation is that we can use them and Hess's Law to calculate the enthalpy change for a reaction that is difficult to measure, or even dangerous. We can look at this as a two step process.

Note the enthalpy of formation is a molar function, so you can have non-integer coefficients. For example, the molar enthalpy of formation of water is:

$$\begin{aligned} H_2(g) + 1/2O_2(g) \rightarrow H_2O(l) \;\; \Delta H^o_f = -285.8\ kJ/mol \\ H_2(g) + 1/2O_2(g) \rightarrow H_2O(g) \;\; \Delta H^o_f = -241.8\ kJ/mol \end{aligned} \tag{10.6.7}$$

This equation says that 85.8 kJ is of energy is exothermically released when one mole of liquid water is formed by reacting one mole of hydrogen gas and 1/2mol oxygen gas (3.011x10^{23} molecules of O_2). Note, ΔH_f^o =of liquid water is less than that of gaseous water, which makes sense as you need to add energy to liquid water to boil it.

Compound	Enthalpy of Formation (kJ/mol)
Hydrogen (H_2)	0
Methane (CH_4)	-74.6
Ethane (C_2H_6)	-84.0
Ethylene (C_2H_4)	52.4
Carbon Dioxide	-393.5
Water (l)	-286
Water (g)	-241.8

Table 10.6.2: Standard enthalpies of formation for select substances. A more comprehensive table can be found at the table of standard enthalpies of formation , which will open in a new window, and was taken from the CRC Handbook of Chemistry and Physics, 84 Edition (2004).

Example 10.6.4: Writing Reaction Equations for ΔH°_f

Write the heat of formation reaction equations for:

a. $C_2H_5OH_{(l)}$
b. $Ca_3(PO_4)_{2(s)}$

Solution

Remembering that ΔH°_f reaction equations are for forming 1 mole of the compound from its constituent elements under standard conditions, we have:

a. $2\,C(s,\ \text{graphite}) + 3\,H_2(g) + \frac{1}{2}O_2(g) \longrightarrow C_2H_5OH(l)$
b. $3\,Ca(s) + \frac{1}{2}P_4(s) + 4\,O_2(g) \longrightarrow Ca_3(PO_4)_2(s)$

Note: The standard state of carbon is graphite, and phosphorus exists as P_4.

Exercise $10.6.3$

Write the heat of formation reaction equations for:

a. $C_2H_5OC_2H_{5(l)}$
b. $Na_2CO_{3(s)}$

Answer a available at the end of this section.

Calculating Enthalpy of Reaction from Standard Enthalpies of Formation

Because enthalpy of reaction is a state function the energy change between reactants and products is independent of the path. We can choose a hypothetical two step path where the atoms in the reactants are broken into the standard state of their element (left side of Figure $10.6.3$), and then from this hypothetical state recombine to form the products (right side of Figure $10.6.3$). Note the first step is the opposite of the process for the standard state enthalpy of formation, and so we can use the negative of those chemical species's $\Delta H_{formation}$.

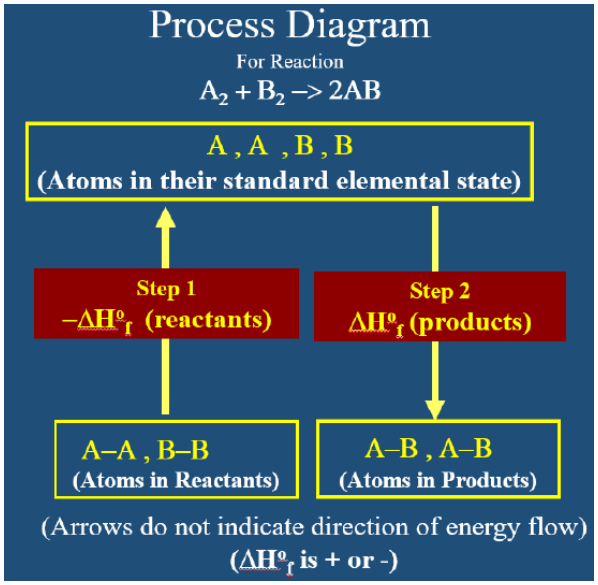


Figure $10.6.3$: Process diagram showing the general process for using enthalpies of formation to calculate enthalpies of reaction. In essence we break apart the reactants into their atoms at standard state (the opposite of formation) in the first step, and then form the products from those atoms in the second step. This is not an energy diagram, but a process diagram, and either step may be endothermic or exothermic.

Note: Enthalpy vs. Molar Enthalpy

Enthalpy, q_p, is an extensive property and for example the energy released in the combustion of two gallons of gasoline is twice that of one gallon. As such, enthalpy has the units of energy (typically J or cal). Molar enthalpies of formation are intensive properties and are the enthalpy per mole, that is the enthalpy change associated with the formation of one mole of a substance from its elements in their standard states. Thus molar enthalpies have units of kJ/mol or kcal/mol, and are tabulated in thermodynamic tables. Also not that the equations associated with molar enthalpies are per mole substance formed, and can thus have non-interger stoichiometric coeffiecents. That is, you can have half a mole (but you can not have half a molecule.

The molar enthalpy of reaction can be used to calculate the enthalpy of reaction if you have a balanced chemical equation. This is described by the following equation, where where m_i and n_i are the stoichiometric coefficients of the products and reactants respectively.

$$\Delta H_{reaction} = \sum m_i \Delta H_f^o(products) - \sum n_i \Delta H_f^o(reactants) \quad (10.6.8)$$

where m_i and n_i are the stoichiometric coefficients of the products and reactants respectively

If we look at the process diagram in Figure 10.6.3 and correlate it to the above equation we see two things.

1. **Reactants break apart** into their atoms that take their elemental standard state forms. This is the opposite of forming them, so this step is the negative of their enthalpy of formation.
2. **Products form** from their atoms in their elemental standard state forms. This is their enthalpy of formation

In both cases you need to multiply by the stoichiomertic coefficients to account for all the species in the balanced chemical equation. This allows us to use thermodynamic tables to calculate the enthalpies of reaction and although the enthalpy of reaction is given in units of energy (J, cal) we need to remember that it is related to the stoichiometric coefficient of each species (review section 5.5.2 enthalpies and chemical reactions).

So the equation

$$\Delta H_{reaction} = \sum m_i \Delta H_f^o(products) - \sum n_i \Delta H_f^o(reactants)$$

describes the enthalpy change as reactants break apart into their stable elemental state at standard conditions and then form new bonds as they create the products.

Let's apply this to the combustion of ethylene (the same problem we used combustion data for).

$C_2H_4 + H_2 \text{ --> } C_2H_6$

using the above equation, we get,

$\Delta H_{reaction} = \Delta H_f^o (C_2H_6) - \Delta H_f^o (C_2H_4) - \Delta H_f^o (H_2)$

ΔH -84 -(52.4) -0= -136.4 kJ

while above we got -136, noting these are correct to the first insignificant digit.

Example 10.6.5

Calculate the enthalpy of formation for acetylene, $C_2H_2(g)$ from the combustion data (table 10.6.1, note acetylene is not on the table) and then compare your answer to the value in table 10.6.2

$\Delta H_{comb} (C_2H_2(g))$ = -1300kJ/mol
$\Delta H_{comb} (C(s))$ = -394kJ/mol
$\Delta H_{comb} (H_2(g))$ = -276kJ/mol

Note, in the following video we used Hess's Law to calculate the enthalpy for the balanced equation, with integer coefficients. It is important that students understand that $\Delta H_{reaction}$ is for the entire equation, so in the case of acetylene, the balanced equation is

$2C_2H_2(g) + 5O_2(g) \text{ --> } 4CO_2(g) + 2 H_2O(l)$ $\Delta H_{reaction} (C_2H_2)$ = -2600kJ

But when tabulating a molar enthaply of combustion, or a molar enthalpy of formation, it is per mole of the species being combusted or formed. That is, the equation in the video and the one above have the exact same value, just one is per mole, the other is per 2 mols of acetylene.

Solution

Watch the video below to get the tips on how to approach this problem.

Step 1: Write out each eq.

$$\begin{aligned} 2C_2H_2(g) + 5O_2(g) &\rightarrow 4CO_2(g) + 2H_2O(l) \quad & \Delta H_{comb} &= -2600kJ \\ C(s) + O_2(g) &\rightarrow CO_2(g) & \Delta H_{comb} &= -393kJ \\ 2H_2(g) + O_2 &\rightarrow 2H_2O(l) & \Delta H_{comb} &= -572kJ \end{aligned} \tag{10.6.9}$$

Step 2: Write out what you want to solve (eq. for the formation of C_2H_2).

$$H_2(g) + 2C(s) \rightarrow C_2H_2(g) \tag{10.6.10}$$

Step 3: Combine given eqs. so they add into desired eq.

- If you reverse Equation change sign of enthalpy
- if you multiply or divide by a number, multiply or divide the enthalpy by that number

$$\begin{aligned} \cancel{2CO_2(g)} + \cancel{H_2O(l)} \rightarrow C_2H_2(g) + \cancel{5/2O_2(g)} \quad & \Delta H_{comb} = -(-\frac{-2600kJ}{2}) \\ 2C(s) + \cancel{2O_2(g)} \rightarrow \cancel{2CO_2(g)} \quad & \Delta H_{comb} = 2(-393kJ) \\ H_2(g) + \cancel{1/2O_2(g)} \rightarrow \cancel{H_2O(l)} \quad & \Delta H_{comb} = \frac{-572kJ}{2} \end{aligned} \tag{10.6.11}$$

Step 4: Sum the Enthalpies: 226kJ (the value in the standard thermodynamic tables is 227kJ, which is the uncertain digit of this number)

See video 10.6.2 for tips and assistance in solving this.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 10.6.2: 6:16 YouTube solving Example 10.6.2 (https://youtu.be/57v2DiU6HjM)

Enthalpies of Reaction and Stoichiometric Problems

Example 10.6.5

Determine the heat released or absorbed when 15.0g Al react with 30.0g Fe_3O_4(s).

Solution

There are three Tasks to this problem

1. Balance Equation and Identify Limiting Reagent
2. Calculate $\Delta H_{reaction}$ from enthalpies of formation
3. Calculate the heat given off by the complete consumption of the limiting reagent

Step 1:

$$\underset{\substack{15.0g\ Al \\ 26.98g/mol}}{8Al(s)} + \underset{\substack{30.0g \\ 231.54g/mol}}{3Fe_3O_4(s)} \rightarrow 4Al_2O_3(s) + 9Fe(3) \tag{10.6.12}$$

$$15gAl\left(\frac{molAl}{26.98g}\right)\left(\frac{1}{8molAl}\right) = 0.069 \tag{10.6.13}$$

$$30.0gFe_3O_4\left(\frac{1molFe_3O_4}{231.54g}\right)\left(\frac{1}{3molFe_3O_4}\right) = 0.043 \tag{10.6.14}$$

So Fe_3O_4(s) is the limiting reagent.

Step 2:

From T1: Standard Thermodynamic Quantities we obtain the enthalpies of formation

ΔH_f^o (Fe_3O_4)	ΔH_f^o (Al_2O_3)	ΔH_f^o (Al)	ΔH_f^o (Fe)
-1118.4kJ/mol	-1675.7kJ/mol	0	0

Using

$$\Delta H_{reaction} = \sum m_i\ \Delta H_f^o \text{ (products)} - \sum n_i\ \Delta H_f^o \text{ (reactants)}$$

gives

$$\Delta H_{reaction} = 4(-1675.7) + 9(0) - 8(0) - 3(-1118.4) = -3363.6kJ$$

Step 3:

Base heat released on complete consumption of limiting reagent.

$$30.0gFe_3O_4\left(\frac{1molFe_3O_4}{231.54g}\right)\left(\frac{-3363kJ}{3molFe_3O_4}\right) = -145kJ \tag{10.6.15}$$

Note, you could have used the 0.043 from step 2,
0.043(-3363kJ)=-145kJ

Example 10.6.7

Using enthalpies of formation from T1: Standard Thermodynamic Quantities calculate the heat released when 1.00 L of ethanol combustion.

Solution

See video 10.6.3

Energy and Stoichiometry.

Pure ethanol has a density of 789g/L. Using the tables for enthalpy of formation, calculate the enthalpy of reaction for the combustion reaction of ethanol, and then calculate the heat released when 1.00 L of pure ethanol combusts.

YouTube Video

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View this YouTube Video online

Video 10.6.3: 4:53min YouTube solving example 10.6.7: Calculating heat released by combustion of 1 L of ethanol (https://youtu.be/57v2DiU6HjM)

Exercise 10.6.4

How much heat is produced by the combustion of 125 g of acetylene?

Answer available at the end of this section.

Contributors and Attributions

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- Ronia Kattoum (UALR)
- Paul Flowers, et al. Open Stax (examples and exercises)
- anonymous

Answers

66.4 kJ

–1407 kJ

$4\,\mathrm{C}(s,\ \text{graphite}) + 5\,\mathrm{H_2}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{C_2H_5OC_2H_5}(l)$;
$2\,\mathrm{Na}(s) + \mathrm{C}(s,\ \text{graphite}) + \frac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{Na_2CO_3}(s)$

6.25×10^3 kJ

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