Guide to the Appendices

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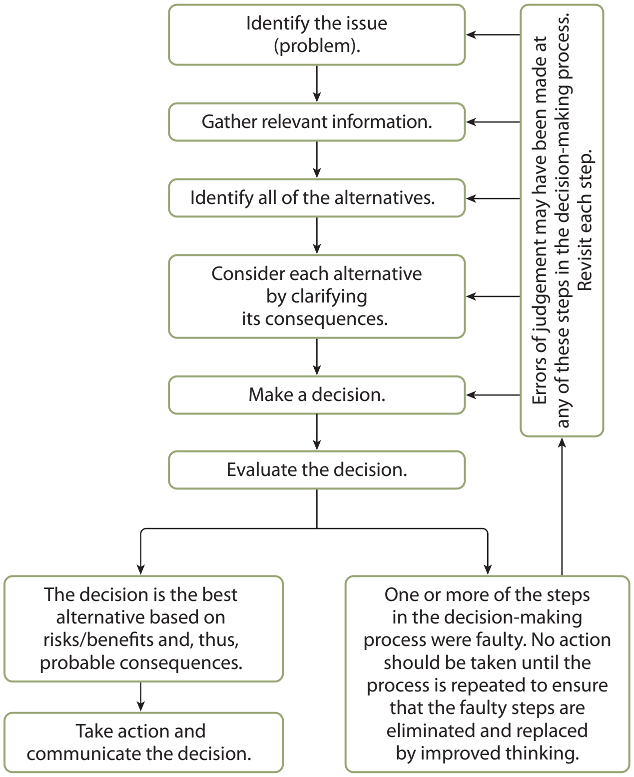
Appendix A Science Skills

Analyzing STSE Issues

STSE is an abbreviation for **s**cience, **t**echnology, **s**ociety, and the **e**nvironment. An issue is a topic that can be seen from more than one point of view. In *Chemistry 11*, you are frequently asked to make connections between scientific, technological, social, and environmental issues. Making such connections could involve, for example, assessing the impact of science on developments in consumer goods, medical devices, or industrial processes; on people, social policy, or the economy; or on air, soil, and water quality, the welfare of organisms, or overall ecosystem health. Analyzing STSE issues involves researching background information about a problem related to science, technology, society, and the environment; evaluating differing points of view concerning the problem; deciding on the best response to the problem; and proposing a course of action to deal with the problem.

The following flowchart outlines one process that can help you to focus your thinking and organize your approach to analyzing STSE issues. The most effective analyses result in decision making and, ultimately, an action plan. Group discussion and collaborative analysis can also play a role in analyzing an STSE issue.

A Process for Analyzing Issues



Identify the Issue (Problem)

An STSE issue is a topic that is debatable—it can be viewed from more than one perspective. When you encounter an issue related to STSE, such as a medical breakthrough, a health-care policy, or an environmental regulation, you need to try to understand it from multiple points of view.

> Suppose you have learned that phosphates from household products, industrial waste, sewage, and agricultural run-off result in algal blooms that reduce biodiversity in rivers, lakes, and oceans.

Assess whether there is any controversy associated with this situation. Could there be different viewpoints concerning the cause of the situation and how to respond to it?

> You read a blog that calls for action on the part of the Canadian government to stop this situation from getting worse. You decide that this situation does represent an STSE issue, because it lends itself to multiple points of view and there is more than one course of action the government could take in response to the situation.

Try to sum up the issue in a specific question.

> For example, "What steps, if any, should the federal government take to reduce phosphates in our waterways?"

Gather Relevant Information

You will need to do some research to gain a better understanding of the issue. Go to Developing Research Skills in Appendix A for help with finding information.

> For example, what role, if any, does the federal government currently play in regulating phosphates? What are the major sources of phosphates that the government regulates? Are there sources of phosphate run-off that are not currently regulated by the government? What methods of regulation could the government impose?

Identify Possible Solutions to the Problem

In order to make an informed decision about how to respond to the issue, you will need to assess the possible solutions to the problem. Your research should reveal some alternative solutions.

> For example, you see a news report in which a political candidate proposes to expand current federal regulations on phosphate levels to include dishwasher detergents rather than just laundry detergents. You read a blog that supports a tax rebate to companies that produce reduced-phosphate fertilizer. One article you read suggests that untreated sewage plays a large role in the high phosphate levels in waterways. Perhaps the government should enact stricter regulations on phosphate levels allowed in waste water from water treatment facilities.

Clarify the Consequences of Each Possible Solution

You may need to do additional research to identify potential consequences of each alternative solution and the reactions of the various stakeholders (that is, the individuals or groups affected by the issue).

> For example:
>
> - How long does it usually take for the government to approve new environmental regulations?
> - Is there a cost associated with enforcing these regulations?
> - Are tax cuts in one category usually replaced with tax increases elsewhere?

You can sort the potential consequences of an alternative into benefits (positive outcomes) and risks (negative outcomes). Use a risk-benefit analysis table like the one below to help you analyze the alternative solutions. For each possible solution, assess the impact on various stakeholders. The potential consequences of each solution could be different for each stakeholder. For some issues, you might choose to assess differing perspectives rather than differing effects on stakeholders. For example, you could assess benefits and risks from economic, environmental, social, scientific, and ethical perspectives. Each perspective could reveal different consequences.

Risk-Benefit Analysis

Issue: What steps, if any, should the federal government take to reduce phosphates in our waterways?

Possible Solutions	Stakeholders	Potential Benefits (positive outcomes)	Potential Risks (negative outcomes)
1. The Canadian government should offer a tax rebate to companies that produce reduced-phosphate fertilizer.	Government	• Reduction in amount of phosphates entering the water system • Less money spent on environmental clean-up	• Reduction in tax revenue • Money would be needed to implement and promote the new rebate
	Fertilizer manufacturer	• Financial incentive to develop a new product • Possible increase in sales due to product being labelled “environmentally friendly” • No job losses, since there is no additional cost to the manufacturer	• Cost of product development may be greater than the rebate offered • Possible loss of revenue due to reduced effectiveness of new product
	Citizen	• Cleaner water leads to healthier, more abundant fish • Healthier ecosystem increases quality of life	• Tax rebate for manufacturers may come at the cost of higher taxes in other categories
	Farmer	• Reduced impact of farming operations on environment and water quality • Effective, environmentally-friendly fertilizer option available	• Possible decrease in crop yield due to reduced effectiveness of new product
2. The Canadian government should develop stricter regulations for sewage treatment.			

Make a Decision

Once you have identified potential outcomes for each possible solution, you are faced with the task of making a decision. Which alternative promises the greatest benefits and the least risks or lowest costs? Your personal values will influence your assessment. You will need to decide whether the benefits of a particular alternative are major or minor. You will also need to decide what an acceptable level of risk is. You might find it helpful to write down a list of questions to help you evaluate the alternative solutions. Some factors to consider are listed here:

- How likely is it that a potential outcome will occur?
- Is there evidence to support the likelihood of a potential outcome?
- How many people (or other organisms) will the proposed course of action affect?
- Is there an estimated sum of money associated with the benefits or costs of each solution?
- Is the outcome of a proposed solution short-term (a one-time benefit/risk) or long-term (ongoing)?
- According to your analysis, how important are the risks of a possible solution compared to its potential benefits?
- How do the benefits and risks of one possible solution compare with the risks and benefits of other possible solutions?

> After considering all the alternatives, you might decide that offering a tax rebate to companies that produce phosphate-free or reduced-phosphate fertilizers will prompt industry change without causing job losses. This solution will have the desired effect of reducing the amount of phosphates entering waterways, while still providing farmers and others with effective fertilizers. Your research also suggests that the cost of providing the rebate is lower than the cost of an environmental clean-up.

Evaluate the Decision

Once you have made a decision, evaluate whether you can justify it with logic and verifiable information. If you discover that some of the information you used to make the decision was incorrect, you should reconsider the alternatives. If new information becomes available, that could also affect your decision.

> Suppose a new study reveals that phosphate levels in ground water continue to rise in a community that banned phosphate fertilizers two years ago. How might this new information affect your decision?

Also, assess whether you have taken all perspectives into account in your analysis. Is there another stakeholder that is strongly affected by a particular alternative? If you decide that you are not confident in the decision you have made, you will need to revisit each step in your analysis.

Act on Your Decision

If you are confident in your decision, the next step is to propose and implement a course of action.

> For example, you could start a community e-mail campaign urging your Member of Parliament to propose that tax rebates be offered to companies that reduce phosphate levels in their products.

Instant Practice

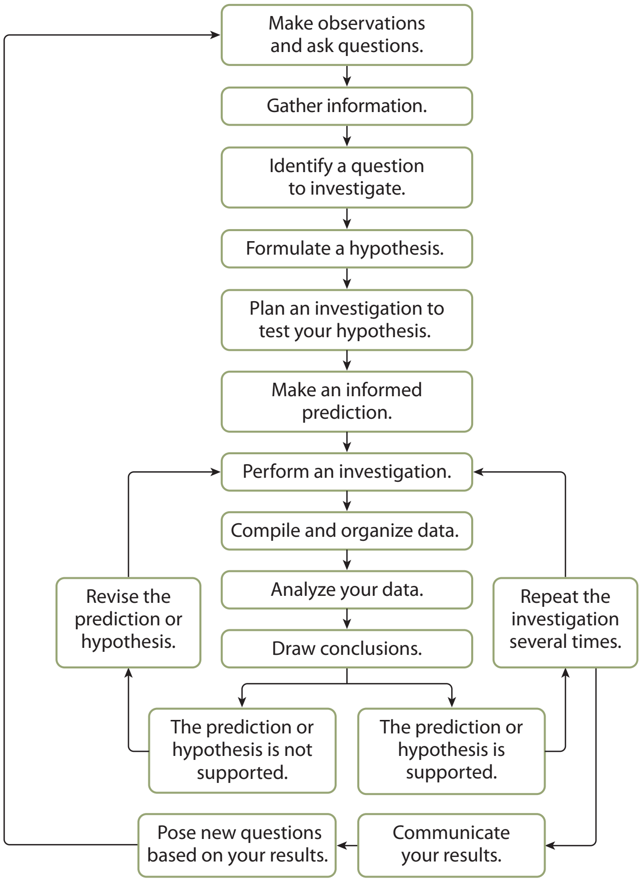
1. Consider the second possible solution listed in the risk-benefit analysis table on the previous page. Create a table in your notes to analyze the benefits and risks of this possible solution. Fill in the "Stakeholders," "Potential Benefits," and "Potential Risks" columns.
2. Look for a chemistry-related STSE issue in the news. Apply the analysis method outlined in this appendix to determine your response to the issue. Write a brief paragraph to explain your viewpoint and a proposed course of action.

Appendix A

Scientific Inquiry

Scientific inquiry is a process that involves making observations, asking questions, performing investigations, and drawing conclusions.

A Process for Scientific Inquiry



Make Observations and Ask Questions

Scientific inquiry usually starts with observations. You notice something that sparks your curiosity and prompts you to ask questions. You try to make sense of your observations by connecting them to your existing knowledge. When your existing knowledge cannot explain your observations, you ask more questions.

For example, suppose a train derailment has resulted in an acid spill near the shore of a local lake. You wonder what effect the spill will have on the fish and other organisms that live there. Has the spill killed all of the organisms in the lake? How have other organisms, such as aquatic plants, been affected by the spill? How can you find answers to your questions without endangering your safety?

Gather Information

Background research may help you to understand your observations and answer some of your questions. Go to Developing Research Skills in Appendix A for guidance on conducting research. You may also be able to gather information by making additional observations.

For example, you read a news report about an environmental assessment of the spill site. You discover that the pH of the lake water before the accident was 6.7. Measurements taken after the accident indicate that the pH dropped to 4.1. You do additional research to find out what kinds of organisms inhabit the lake and the optimal pH for their survival and growth.

Identify a Question to Investigate

You need to have a clear purpose and decide on a specific question that you are able to investigate with the resources available. If a question is provided for you, make sure you understand the science behind the question.

You decide to investigate the effect of acidity on living organisms. You do not wish to risk harming fish or other animals, so you decide to use aquatic plants as your test organism. You pose the scientific question, “What effect will increasing acidity have on aquatic plants grown in an aquarium?”

Formulate a Hypothesis

A hypothesis attempts to answer the question being investigated. It often proposes a relationship that is based on background information or an observed pattern of events.

You hypothesize that because plants can remove some impurities from polluted water, aquatic plants will be able to reduce the effect of small amounts of acid. However, because highly acidic water will damage or kill most organisms, you hypothesize that the aquatic plants will not be able to counteract the addition of large quantities of acid.

Plan an Investigation

Some investigations lay out steps for you to follow in order to answer a question, analyze a set of data, explore an issue, or solve a problem. In planning your own investigation, however, *you* must decide how to approach a scientific question. Taking time to plan your approach thoroughly will ensure that you address the question appropriately.

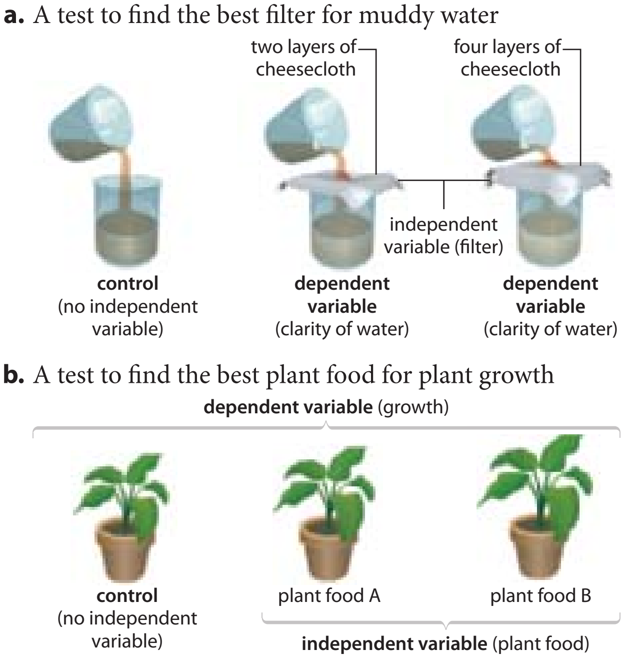
Design a Procedure Write out step-by-step instructions for performing the investigation. Include instructions for repeat trials, if appropriate. Ensure that the procedure is written in a logical sequence, and that it is complete and clear enough that someone else could carry it out. Create diagrams, if necessary. Ask someone else to read through the procedure and explain it back to you, to ensure you have not omitted any important details.

You decide to investigate the change in the pH of water when you add acid to a large glass bottle containing water and aquatic plants. You will measure the pH of the water and observe the physical appearance of the plants twice a day for three days.

Identify Variables Many investigations study relationships between variables (quantities or factors that can change). An *independent variable* is changed by the person conducting the investigation. A *dependent variable* is affected by changes in the independent variable. *Controlled variables* are kept the same throughout an experiment.

A simple, controlled experiment shows relationships especially clearly because it has a single independent variable and a single dependent variable. All other variables are controlled. Changes in the dependent variable occur only in response to changes in the independent variable. When you are planning your investigation, you will need to identify the variables and decide which ones to control.

If possible, investigations include a *control:* a situation identical to the one being tested, except that the independent variable is not changed in any way. There is no reason, therefore, for the dependent variable to change. If it does, the reasoning behind your hypothesis, prediction, and variable analysis may be faulty. Look at the illustration at the top of the next column to see some examples of independent and dependent variables, as well as two examples of a control (no independent variable).



In planning your investigation, you decide to manipulate (change) the quantity of acid added to the aquarium you have made. Therefore, the quantity of acid is the independent variable. The pH of the water will be the dependent variable. Water temperature, lighting conditions, the particular species of aquatic plant used, and nutrients will be the controlled variables. In addition, you decide to set up an identical aquarium as a control. No acid will be added to the water in this aquarium, so you expect the pH of the water to stay constant.

You decide to set up three different experimental aquariums. You plan to add a different amount of acid to the water in each of these aquariums. You will add no acid to the fourth aquarium (the control). Your teacher suggests using 1.0 mol/L hydrochloric acid. You will add five drops of hydrochloric acid to Aquarium 1, 10 drops to Aquarium 2, and 15 drops to Aquarium 3. You will measure the pH of the water in each aquarium at 9:00 A.M. on the first day, immediately before adding the acid, and then at 2:00 P.M. You will measure the pH again at 9:00 A.M. and 2:00 P.M. on Day 2 and Day 3.

List Materials and Safety Precautions Develop a list of materials and apparatus you will need. Include measuring and recording instruments. Examine your procedure for safety hazards and plan any necessary precautions. (Go to Safety in Your Chemistry Lab and Classroom at the front of this book for information about safety hazards and precautions.) **Note:** Before doing any experimental work, ask your teacher to examine and approve your plan.

Your materials list will include safety goggles, a lab coat or apron, protective gloves, four glass jars, a dropper, aquatic plants, hydrochloric acid, water, and a pH meter. Safety precautions include handling glassware carefully to avoid breakage; wearing safety eyewear, gloves, and protective clothing to protect yourself from any acid spillage; storing the acid safely after use; and disposing of the aquarium water at the end of the investigation according to your teacher's instructions.

Make an Informed Prediction

A clear hypothesis often leads to a specific, testable prediction about what the investigation will reveal. You need to determine how to test your question before you can predict what will happen.

You predict that an aquarium full of a certain species of aquatic plants will maintain a stable pH of about 7 when a small quantity of acid is added. When greater quantities of acid are added, however, you predict that the plants will be damaged. The pH of the water will decrease rapidly and the plants will eventually die.

Perform an Investigation

Be responsible whenever you conduct an investigation. Think before acting, and follow all safety precautions. Carry out your procedure carefully. Ask for assistance if you are unsure how to proceed or if you encounter an unexpected difficulty. Report any accidents to your teacher immediately. Keep your workspace neat and clean it up when you have finished your investigation.

Compile and Organize Data

Record your results carefully and organize them in a logical way. Go to Organizing Data in a Table in Appendix A for help with recording and organizing the results of an investigation. As part of your observations, keep careful notes of any unexpected occurrences, problems with equipment, or unusual circumstances that might affect your results. If you are working with a partner, ensure that both of you have a copy of all observations and results.

Your results may include either qualitative or quantitative observations, or both. *Quantitative observations* are measurable and involve numbers. *Qualitative observations* involve descriptions rather than numbers or measurements. When making qualitative observations, try to record specific characteristics so that you can make comparisons between different trials.

In your investigation, you will record both qualitative and quantitative results. The pH values that you record are quantitative observations. Your descriptions of the physical appearance of the aquatic plants are qualitative observations. Looking at specific plant characteristics such as colour (green or brown) and vigour (robust or spindly) will help you to compare the physical appearance of the plants in each aquarium.

You might use a table like the one below to record and organize the data from your investigation.

Effect of Increasing Acidity on the Physical Appearance of Aquatic Plants in an Aquarium

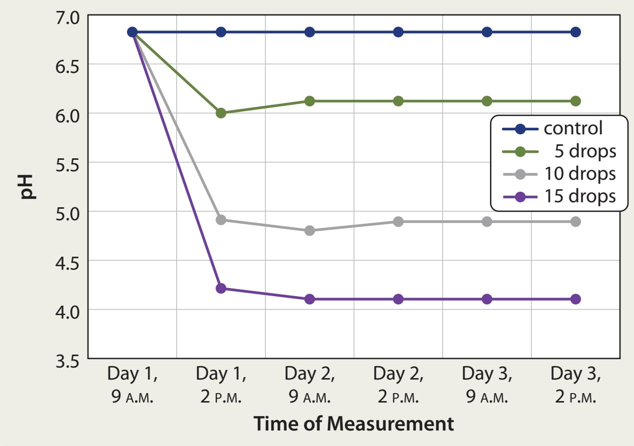
	Physical Appearance of Plants (colour and vigour)			
	Control	+ 5 drops of acid	+ 10 drops of acid	+ 15 drops of acid
Day 1, 9:00 A.M. (before addition of acid)	green, robust	green, robust	green, robust	green, robust
Day 1, 2:00 P.M.	green, robust	green, robust	green, robust	brownish spots, spindly
Day 2, 9:00 A.M.	green, robust	green, robust	brownish spots, less robust	brown, spindly (looks dead)
Day 2, 2:00 P.M.	green, robust	green, robust	mostly brown, less robust	brown, spindly (looks dead)
Day 3, 9:00 A.M.	green, robust	green, robust	brown, spindly (looks dead)	brown, spindly (looks dead)
Day 3, 2:00 P.M.	green, robust	green, robust	brown, spindly (looks dead)	brown, spindly (looks dead)

Analyze Your Data

Perform any necessary graph work or calculations. Go to Constructing Graphs in Appendix A for help with graphing. Then consider and interpret your results. Do your data and observations support or refute your hypothesis and prediction? Are additional data needed before you can draw definite conclusions? Identify any possible sources of error or bias in your investigation. Does the procedure or apparatus need to be modified to obtain better data?

Using the pH data from your investigation, you construct the graph shown below.

Effect of Increasing Acidity on pH in an Aquarium



Draw Conclusions

Conclusions usually answer several questions:

- What has the investigation revealed about the answer to the question?
- How well does your prediction agree with the data?
- How well is your hypothesis supported by the data? Are the observations explained by the hypothesis?
- How precise were the measuring instruments and resulting observations?
- What improvements could be made to the investigation?

Relate your conclusions to your background knowledge of the scientific principles involved.

You conclude that your hypothesis and prediction were supported by the data. The aquatic plants were able to survive the addition of small amounts of hydrochloric acid (5 drops), but when more acid was added to the aquarium, the pH decreased rapidly and the plants soon looked brown and spindly.

However, you are unsure whether the plants in the highly acidic water were actually dead. You think perhaps you could place them in fresh water to see whether they could recover.

Now you can relate your results to the original problem of the acid spill in the lake. Are the aquatic plants in the lake likely to survive? Do your results reflect the conditions in the lake? What other factors might you need to consider?

Communicate Your Results

Communicate the results of your investigation. Always include a summary of your findings and an evaluation of the investigation. Be sure to round answers to the proper number of significant digits. Go to Significant Digits in Appendix A for help with reporting numerical answers. Demonstrate your results clearly using graphs, tables, or diagrams, as appropriate. Go to Constructing Graphs or Organizing Data in a Table in Appendix A for help with communicating your results. Be sure to include units when expressing measurements. Go to Measurement in Appendix A for information on units and measurements.

Pose New Questions Based on Your Results

The conclusion of an investigation is not the end of scientific inquiry. Scientific inquiry is a continuous process in which results and conclusions lead to new questions. What new research questions might arise from your investigation? How might you find an answer to one of these questions?

After performing this investigation, you wonder how much the aquatic plants themselves affected the pH of the water. Would the decrease in pH have been more noticeable if there were no plants present? How would you test this?

Instant Practice

You are asked to plan an investigation to determine whether calcium chloride or sodium chloride is a better de-icing agent. Think about how you could test a hypothesis related to this question.

1. Will your results include qualitative or quantitative observations, or both? Explain.
2. State a hypothesis for this investigation.
3. What will your independent variable be? What will your dependent variable be? What control will you set up?

Appendix A

Developing Research Skills

In this course, you will need to conduct research to answer specific questions and to explore broad research topics. The following skills will take you through the research process from start to finish:

- focussing your research
- searching for resources that contain information related to your topic
- evaluating the reliability of your information sources
- gathering, recording, and organizing information in an appropriate format
- presenting your work

Focussing Your Research

- Start by carefully reading your assignment. Pick out key words and phrases, such as *apply*, *analyze*, *argue*, *compare and contrast*, *describe*, *discuss*, *evaluate*, *explain*, *identify*, *infer*, *interpret*, and *predict*. These key words and phrases will guide you on what kind of information you need to collect, and what you need to do with the information.
- Jot down ideas on your own, and then get additional input from others, including your teacher.
- Once you have done some general research, narrow down your topic until you can express it in one specific question. This will help you focus your research.
- Ensure that the question you are researching fulfills the guidelines of the assignment provided by your teacher.

Searching for Resources

- It is important to find reliable resources to help you answer your question. Potential sources of information include print and on-line resources such as encyclopedias, textbooks, non-fiction books, journals, websites, and newsgroups.
- The library and the Internet can both provide information for your search. Whether you are looking at print or digital resources, you need to evaluate the accuracy and objectivity of the information.

Evaluating the Reliability of Your Information Sources

Assess the reliability of your information sources to help you decide whether the information you find is likely to be accurate. To determine the validity of a source, check that the author is identified, a recent publication date is given, and the source of facts or quotations is identified. An author's credentials are important. Look for an indication of educational background, work experience, or professional affiliation. If the information is published by a group, try to find out what interests the group represents. The following guidelines may be helpful in assessing your information sources:

- On-line and print scientific journals provide data that have been reviewed by experts in a field of study (peer-reviewed), so they are usually a reliable source. Be aware, however, that the conclusions in journal articles may contain opinions as well as facts.
- Data on the websites of government statistical departments tend to be reliable. Be sure to read carefully, however, to interpret the data correctly.
- University resources, such as websites ending in ".edu" are generally reliable.
- Reliable experts in a field of study often have a PhD or MSc degree, and their work is regularly cited in other publications.
- Consumer and corporate sources may present a biased view. That is, they may only present data that support their side of an issue. Look for sources that treat all sides of an issue equally and fairly, or that clearly specify which perspective(s) they are presenting.
- Some sources, such as blogs and editorials, provide information that represents an individual's point of view or opinion. Therefore, the information is not objective. However, opinion pieces can alert you to controversy about an issue and help you consider various perspectives. The opinion of an expert in a field of study should carry more weight than that of an unidentified source.
- On-line videos and podcasts can be dynamic and valuable sources of information. However, their accuracy and objectivity must be evaluated just as thoroughly as all other sources.

- A piece of information is generally reliable if you can find it in two other sources. However, be aware that several on-line resources might use the same incorrect source of information. If you see identical wording on multiple sites, try to find a different source to verify the information.

Gathering, Recording, and Organizing Information

- As you locate information, you may find it useful to jot it down on large sticky notes or make colour-coded entries in a digital file so you can group similar ideas together. Remember to document the source of your information for each note or data entry.

Avoid Plagiarism Copying information word-for-word and then presenting it as your own work is called *plagiarism*. Instead, you must cite every source you use for a research assignment. This includes all ideas, information, data, and opinions, other than your own, that appear in your work. If you include a quotation, be sure to indicate it as such, and supply all source information. Avoid direct quotations whenever possible—put information in your own words. Remember, though, that even when you paraphrase, you need to cite your sources.

Record Source Information A research paper should always include a bibliography—a list of relevant information sources you have consulted while writing the paper. Bibliographic entries include information such as the author, title, publication year, name of the publisher, and city in which the publisher is located. For magazine or journal articles, the name of the magazine or journal, the name of the article, the issue number, and the page numbers should be recorded. For on-line resources, you should record the site URL, the name of the site, the author or publishing organization, and the date on which you retrieved the information. Remember to record source information while you are taking notes to avoid having to search it out again later! Ask your teacher about the preferred style for your references.

- You might find it helpful to create a chart to keep track of detailed source information. For on-line searches, a tracking chart is useful to record the key words you searched, the information you found, and the URL of the website where you found the information.
- Write down any additional questions that you think of as you are researching. You may need to refine your topic if it is too broad, or take a different approach if there is not enough information available to answer your research question.

Presenting Your Work

- Once you have organized all of your information, you should be able to summarize your research so that it provides a concise answer to your original research question. If you cannot answer this question, you may need to refine the question or do a bit more research.
- Check the assignment guidelines for instructions on how to format your work.
- Be sure that you fulfill all of the criteria of the assignment when you communicate your findings.

Instant Practice

1. Your assignment asks you to research "green" cleaners and present your opinion on which are the least harmful to the environment and why.
   a. What search terms might you use for your initial research on the Internet or at the library?
   b. How might you narrow down this assignment into a research question?
2. How could you evaluate the reliability of an on-line video about air pollution?
3. Wiki sites allow users to contribute and edit content.
   a. How could this affect the reliability of the information they present?
   b. What steps would you take to verify a piece of information you found in a wiki entry?

Appendix A

Writing a Lab Report

Use the following headings and guidelines to create a neat and legible lab report.

Title

- Choose a title that clearly states the independent variable and the dependent variable, but not the outcome of the investigation. For example, "A Comparison of the Neutralizing Ability of Different Antacid Ingredients."
- Under the title, write the names of all participants and the date(s) of the investigation.

Introduction

- Summarize the background of the problem.
- Cite any relevant scientific principles or literature related to the question being investigated.

Question/Problem

- Clearly state the question being investigated or the problem for which you are seeking a solution. For example, "Which ingredient in antacids is most effective at neutralizing acid?"

Hypothesis

- State, in general terms, the relationship that you believe exists between the independent variable and the dependent variable. For example, "Calcium carbonate, an ingredient in many antacids, is more effective at neutralizing acid than sodium hydrogen carbonate, another ingredient found in antacids."

Prediction

- State, in detailed terms, the specific results you expect to observe. For example, "Calcium carbonate will neutralize more hydrochloric acid than the same mass of sodium hydrogen carbonate will."

Materials

- List all of the materials and equipment you used, or refer to the appropriate page number in your textbook, and note any additions, deletions, or substitutions you have made.

Procedure

- Write your procedure in the form of precise, numbered steps, or refer to the appropriate page number in your textbook, and note any changes to the procedure. Include any safety precautions.

Results

- Set out the observations and/or data in a clearly organized table(s). Give your table(s) a title.
- If appropriate, construct a graph that shows the data accurately. Label the x-axis and the y-axis of the graph clearly and accurately, and use the correct scale and units. Give your graph a title.

Data Analysis

- Analyze all the results you have gathered and recorded, and ensure that you can defend your analysis. For example, "As shown in the following calculations, the volume of hydrochloric acid neutralized by calcium carbonate was 1.5 mL more on average than that neutralized by sodium hydrogen carbonate."
- Show sample calculations for any mathematical data analysis.

Conclusion

- State a conclusion based on your data analysis. Relate your conclusion to your hypothesis. For example, "Based on the results of this investigation, calcium carbonate has a greater neutralizing ability than sodium hydrogen carbonate."
- Compare the results you obtained with those you expected, or those obtained by other researchers.
- Examine and comment on experimental error.
- Assess the effectiveness of the experimental design.
- Indicate how the data support your conclusion.
- Make recommendations for how your conclusion could be applied, or for further study of the question you investigated.

References

- Cite your information sources according to the reference style your teacher suggests.
- Sources that need to be cited include background information for your introduction, a materials list or procedure from a textbook, any specialized methods of data analysis, results from other studies that you used for comparison with your own results, and any other sources used in your conclusion.

Appendix A

Organizing Data in a Table

Scientific investigation is about collecting information to help you answer a question. In many cases, you will develop a hypothesis and collect data to see if your hypothesis is supported. An important part of any successful investigation is recording and organizing your data. Often, scientists create tables in which to record data.

Planning to Record Your Data Suppose you are doing an investigation on the water quality of a stream that runs near your school. You will take water samples at three different locations along the stream. You need to decide how to record and organize your data. Begin by making a list of what you need to record. For this experiment, you will need to record the sample site, the pH of the water at each sample site, the chemicals found in the water at each sample site, and the concentration of these chemicals.

Creating Your Data Table Your data table must allow you to record your data neatly. To do this you need to create

- headings to show what you are recording
- columns and rows that you will fill with data
- enough cells to record all the data
- a title for the table

In this investigation, you will find several chemicals in the water at each site, so you must make space for multiple recordings at each site. This means every row representing a sample site will have at least four rows associated with it for the different chemicals.

If you think you might need extra space, create a special section. In this investigation, leave space at the bottom of your table, in case you find more than four chemicals in the water at a sample site. Remember, if you use the extra rows, make sure you identify which sample site the extra data are from. Finally, give your table an appropriate title. Your data table might look like the one in the next column.

Reading a Table A table can be used to organize observations and measurements so that data are represented neatly and clearly. However, a table can also show relationships among the data presented. When you are reading a table, be sure to start by reading the column and row headings carefully. If the table contains measurements, look for the units in which they are reported. Follow vertically down a column or horizontally across a row to look for trends in the data. If the table contains numbers, do the numbers increase or decrease as you look down the column or across the row?

Water Quality Observations Made at Three Sample Stream Sites

headings show what is being recorded

columns and rows contain data

Sample Site	pH	Type of Chemical	Concentration of Chemical (mg/L)
1	6.9	sulfate	30.7
		nitrate	0.11
		phosphate	0.001
2	7.2	sulfate	31.2
		nitrate	0.35
		phosphate	0.002
3	7.1	sulfate	30.9
		nitrate	0.07
		phosphate	0.001
		iron	0.1

extra rows to collect data in case you need to add observations

Also look for relationships between columns or rows. Do the numbers in one column increase as the numbers in another column decrease? Is there one piece of data that does not fit the pattern in the rest of the table? Think about why this might be the case.

Instant Practice

1. You want to compare the antibiotic effects of silver (found in wound dressings) and penicillin. Construct a table to record the number of Bacteria A, Bacteria B, and Bacteria C growing on three different media:
   - a standard culture medium
   - a standard medium with penicillin added
   - a standard medium with silver added
2. Now you wish to refine your investigation to record the number of bacteria of each type growing on each medium after 12 hours, 24 hours, and 36 hours. Draw a new table to record these data.
3. Examine the table at the top of the page. What does it tell you about the three sample sites?

Appendix A

Constructing Graphs

A graph is a diagram that shows relationships among variables. Graphs help you to interpret and analyze data. The three basic types of graphs used in science are the line graph, the bar graph, and the circle graph.

The instructions given here describe how to construct graphs using paper and pencil. You can also use computer software to generate graphs. Whichever method you use, the graphs you construct should have the features described in the following pages.

Line Graphs

A line graph is used to show the relationship between two variables. The independent variable is plotted on the horizontal axis, called the *x*-axis. The dependent variable is plotted on the vertical axis, called the *y*-axis. The dependent variable (*y*) changes as a result of a change in the independent variable (*x*).

Suppose a chemist carried out an investigation to determine the relationship between the temperature and volume of an unknown gas at a specific pressure (P_1). She measured the volume (in mL) of the gas upon heating it to various temperatures (in K), as shown in the table below.

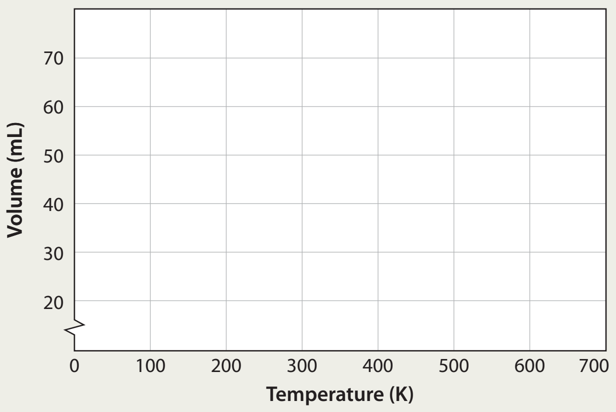
Volume and Temperature for an Unknown Gas at P_1

Volume (mL)	Temperature (K)
38	300
49	400
62	500
75	600

To make a graph of volume versus temperature measurements for this gas, start by determining the dependent and independent variables. The volume of the gas is the dependent variable and is plotted on the *y*-axis. The independent variable, or the temperature to which the the gas was heated, is plotted on the *x*-axis.

Give your graph a title and label each axis, indicating the units if appropriate. In this example, label the temperature on the *x*-axis. Your *x*-axis will need to be numbered to at least 600 K. Because the lowest volume of gas measured was 38 mL and the highest was 75 mL, you know that you will have to start numbers on the *y*-axis from at least 38 and number to at least 75 mL. For instance, you could decide to number 20 to 80 by intervals of 10, spaced at equal distances. Look at the example at the top of the page to see how you could label your axes.

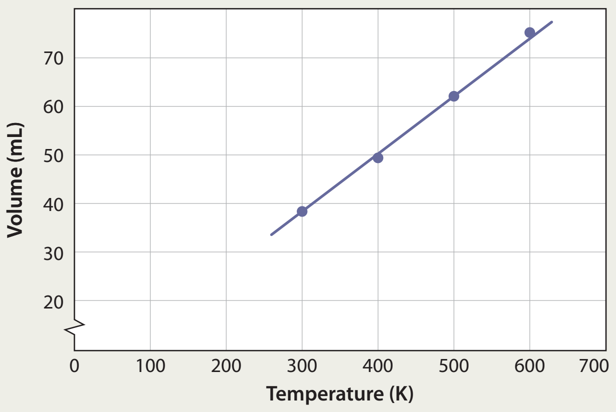
The Relationship between Volume and Temperature for an Unknown Gas at P_1



Begin plotting points by locating 300 on the *x*-axis and 38 on the *y*-axis. Where an imaginary vertical line from the x-axis and an imaginary horizontal line from the *y*-axis meet, place the first data point. Place other data points using the same process. After all the points are plotted, draw a "best fit" straight line through the points.

A best fit line should be drawn to represent the general trend of the data. Try to draw the line so that there are as many points above it as there are below. Do not change the position or slope of the line dramatically just to include an outlier—a single data point that does not seem to be in line with all the others.

The Relationship between Volume and Temperature for an Unknown Gas at P_1



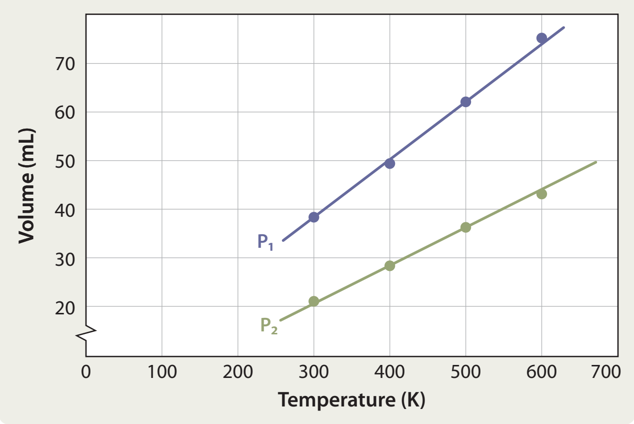
The chemist then repeated the investigation at a different pressure (P_2), using the same amount of gas. Her observations are shown in the table at the top of the next page.

Volume and Temperature for an Unknown Gas at P_2

Volume (mL)	Temperature (K)
21	300
28	400
36	500
43	600

What if you want to compare the relationship between volume and temperature of the gas at these two different pressures? The P_2 data can be plotted on the same graph as the data for P_1. Label the different lines indicating different sets of data as P_1 and P_2.

The Relationship between Volume and Temperature at P_1 and P_2



Instant Practice

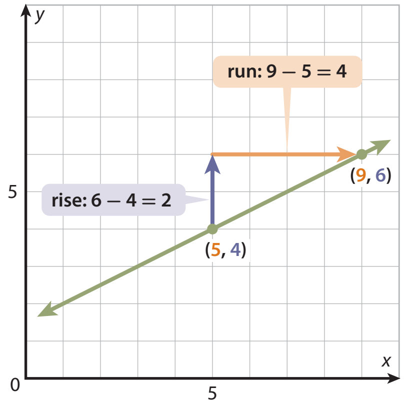
1. Between 400 K and 500 K, what change in volume occurred at P_1?
2. As temperature increased, how did the changes in volume at P_1 compare to the changes in volume at P?
3. Construct a line graph for the following data:

Partial Pressures of Water Vapour at Different Temperatures

Temperature (°C)	Pressure (kPa)
15	1.71
16	1.81
17	1.93
18	2.07
19	2.20

Slope of a Linear Graph The slope of a line is a number determined by any two points on the line. This number describes how steep the line is. The greater the absolute value of the slope, the steeper the line. Slope is the ratio of the change in the *y*-coordinates (rise) to the change in the *x*-coordinates (run) as you move from one point to the other.

The graph below shows a line that passes through points (5, 4) and (9, 6).



$$\text{Slope} = \frac{\text{rise}}{\text{run}}$$

$$= \frac{\text{change in } y\text{-coordinates}}{\text{change in } x\text{-coordinates}}$$

$$= \frac{6 - 4}{9 - 5}$$

$$= \frac{2}{4} \quad \text{or} \quad \frac{1}{2}$$

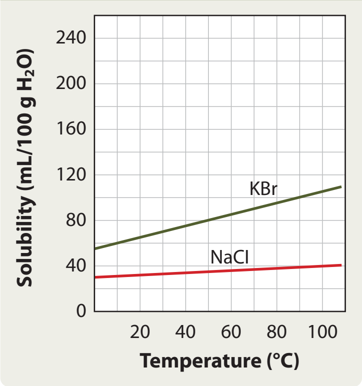
So, the slope of the line is $\frac{1}{2}$.

A positive slope indicates that the line climbs from left to right. A negative slope indicates that the line descends from left to right. A slope of zero indicates that there is no change in the dependent variable as the independent variable increases. A horizontal line has a slope of zero.

Linear and Non-Linear Trends Two types of trends you are likely to see when you graph data in chemistry are linear trends and non-linear trends. A linear trend has a constant increase or decrease in data values. For a non-linear trend, the degree to which the data values are increasing or decreasing is not constant. The graphs shown on the next page are examples of these two common trends.

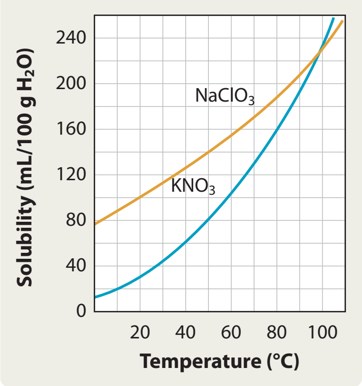
In the graph below, there are two lines describing the solubility of salts at various temperatures. Both lines show an increasing, linear trend. As the temperature increases, so does the solubility of each salt. The rate of increase is constant.

Solubility of Two Salts at Various Temperatures



The graph below also shows two lines describing the solubility of salts at various temperatures. Both lines show an increasing, non-linear trend. As in the graph above, the solubility of each salt increases as the temperature increases. However, for the graph below, the rate of increase is not constant. For instance, for potassium nitrate, you will see that the compound's solubility increases more as the temperature increases 20°C from 60°C to 80°C than it does as it increases 20°C from 30°C to 50°C.

Solubility of Two Salts at Various Temperatures

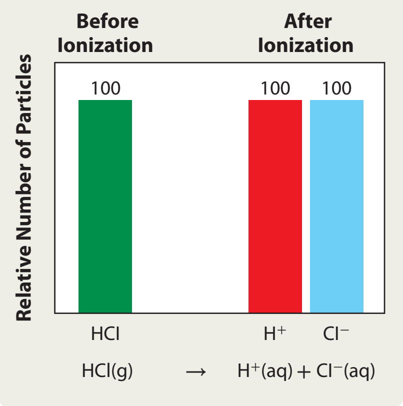


When you are drawing a curve to represent a non-linear trend, you should not connect the data points. Instead, draw a smooth best-fit curve that shows the general trend of the data. Try to draw the curve so there are as many points above it as there are below. The curve should change smoothly. It should not have a dramatic change in direction just to include a single data point that does not fit with the others.

Bar Graphs

A bar graph displays a comparison of different categories of data by representing each category with a bar. The length of the bar is related to the category's frequency. To make a bar graph, set up the *x*-axis and *y*-axis as you did for the line graph. Plot the data by drawing thick bars from the *x*-axis up to an imaginary line representing the *y*-axis point.

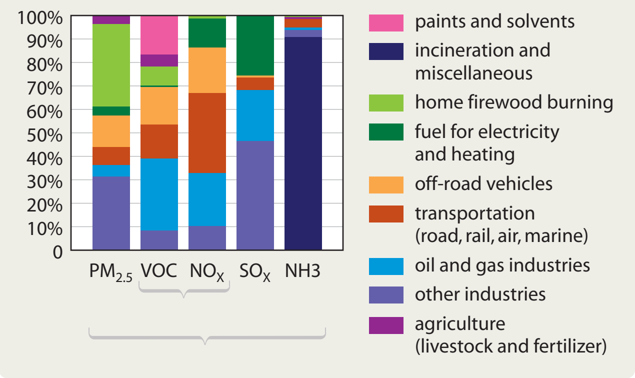
Ionization of a Strong Acid in Water



Look at the graph above. The independent variable is the type of particle. The dependent variable is the relative number of particles.

Bar graphs can also be used to display multiple sets of data in different categories at the same time, as shown in the the graph below. Bar graphs like the one below have a legend to denote which bars represent each set of data.

Emission Sources of Air Pollutants Leading to Ground-Level Ozone and Fine Particulate Matter, 2006



Instant Practice

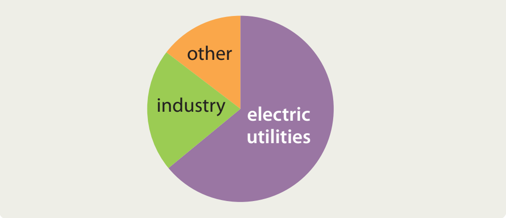
In the graph at the top of this column, what is the relative number of chlorine ions after ionization?

Circle Graphs

A circle graph is a circle divided into sections that represent parts of a whole. When all the sections are placed together, they equal 100 percent of the whole.

Consider the circle graph shown below. This graph shows the anthropogenic sources of the air contaminant sulfur dioxide. Each component of the graph, electrical utilities, industry, and other, add up to 100 percent of all sources of sulfur dioxide air pollution.

Anthropogenic Sources of Sulfur Dioxide



Suppose you wanted to make a circle graph to represent data you observed or calculated, such as the percentage composition of a compound. For instance, if you determine the percentage composition of copper(I) sulfide, $Cu_2S(s)$, to be 79.9% copper and 20.1% sulfur by mass, you can represent this graphically with a circle graph.

To begin, you know that the percent of the different elements in the compound must add up to 100. This 100 percent is represented by the 360° (the number of degrees in a circle) that make up the circle graph.

To find out how much of the circle each element should cover in the graph, first multiply the percent of copper by 360. Then, round your answer to the nearest whole number.

$$\begin{aligned} 79.9\% \times 360° &= 0.799 \times 360° \\ &= 287.64° \\ &= 287° \end{aligned}$$

The sum of all the segments of the circle graph should add up to 360°. Therefore, you can calculate the segment of the circle that represents the percent of sulfur by subtracting the degrees representing copper from 360°.

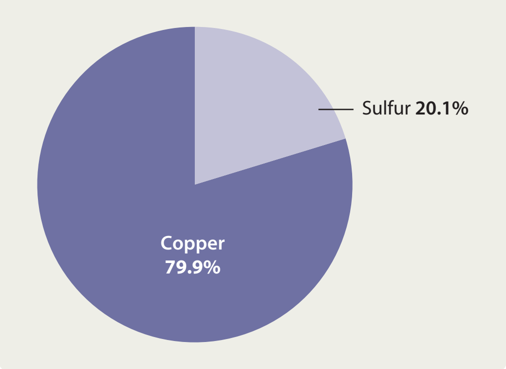
$$360° - 287° = 73°$$

To draw your circle graph, you will need a compass and a protractor. First, use the compass to draw a circle. Then, draw a straight line from the centre to the edge of the circle. Place your protractor on this line, and mark the point on the circle where an angle of 73° will intersect the circle. Draw a straight line from the centre of the circle to the intersection point. This is the section representing the percent of sulfur in the compound. The remaining section represents the percent of copper.

Complete the graph by labelling the sections of the graph with percentages and giving the graph a title. Your completed graph should look similar to the one below.

If your circle graph has more than two sections, you will need to construct a segment for each entry. Place your protractor on the last line segment that you have drawn and mark off the appropriate angle. Draw a line segment from the centre of the circle to the new mark on the circle. Continue this process until all of the segments have been drawn.

Percentage Composition of Copper(I) Sulfide



Instant Practice

Create a circle graph to illustrate the percentage composition of the mineral chalcopyrite, $CuFeS_2(s)$. Based on your calculations, you know that the percentage composition of chalcopyrite is 34.62% copper, 30.43% iron, and 34.95% sulfur, by mass.

Appendix A

Using Graphic Organizers

When deciding which type of graphic organizer to use, consider your purpose. It may be to brainstorm, to show relationships among ideas, to summarize a section of text, to record research notes, or to review what you have learned before writing a test. Several different graphic organizers are shown here. The descriptions indicate the function or purpose of each organizer.

PMI Chart

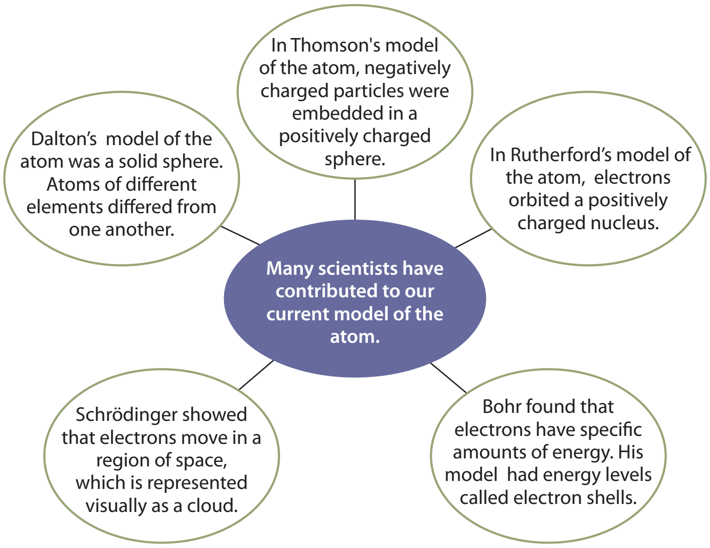
PMI stands for Plus, Minus, and Interesting. A PMI chart is a simple three-column table that can be used to state the positive and negative aspects of an issue, or to describe advantages and disadvantages related to the issue. The third column in the chart is used to list interesting information related to the issue. PMI charts help you organize your thinking after reading about a topic that is up for debate or that can have positive or negative effects. They are useful when analyzing an issue.

Base metal smelting is an important industry in Canada. However, chemicals released during smelting can endanger people's health.

P	M	I
Base metal smelting produces useful metals, making it an important industry in Canada.	Harmful chemicals can be released into the environment during smelting, endangering the health and safety of local populations.	Many of the substances emitted by base metal smelters are listed as toxic by the Canadian Environmental Protection Act.
The smelting industry employs thousands of Canadians and contributes billions of dollars to the Canadian economy.	Smelting companies argue that setting strict limits on smelting emissions would place them at a competitive disadvantage, which would affect jobs.	Many smelters are located in remote areas. Finding other types of employment in these areas can be challenging.

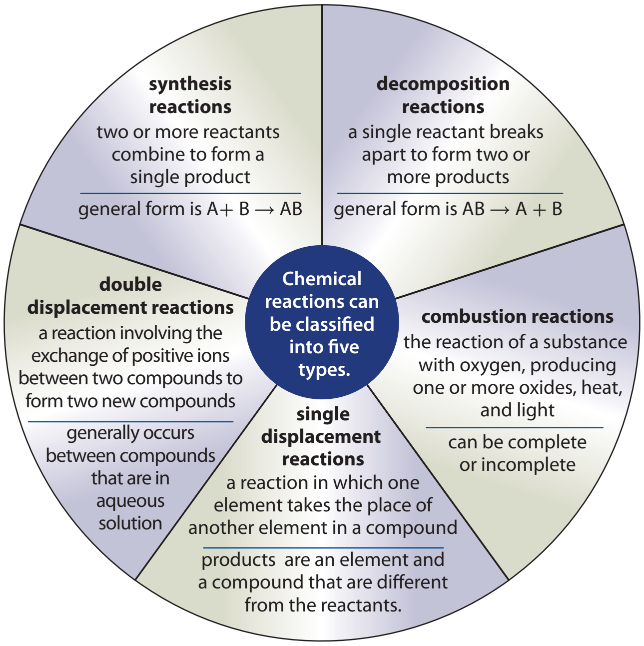
Main Idea Web

A main idea web shows a main idea and several supporting details. The main idea is written in the centre of the web, and each detail is written at the end of a line extending from the centre. This organizer is useful for brainstorming or for summarizing text.



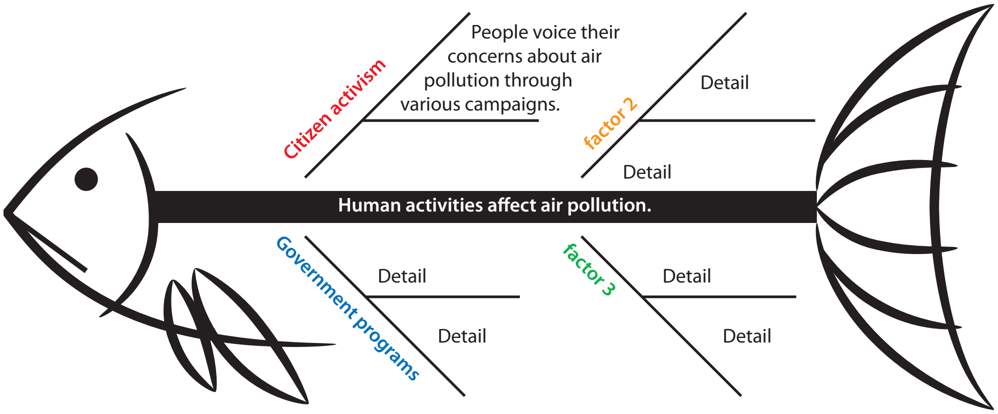
Spider Map

A spider map shows a main idea and several ideas associated with the main idea. It does not show the relationships among the ideas. A spider map is useful when you are brainstorming or taking notes.



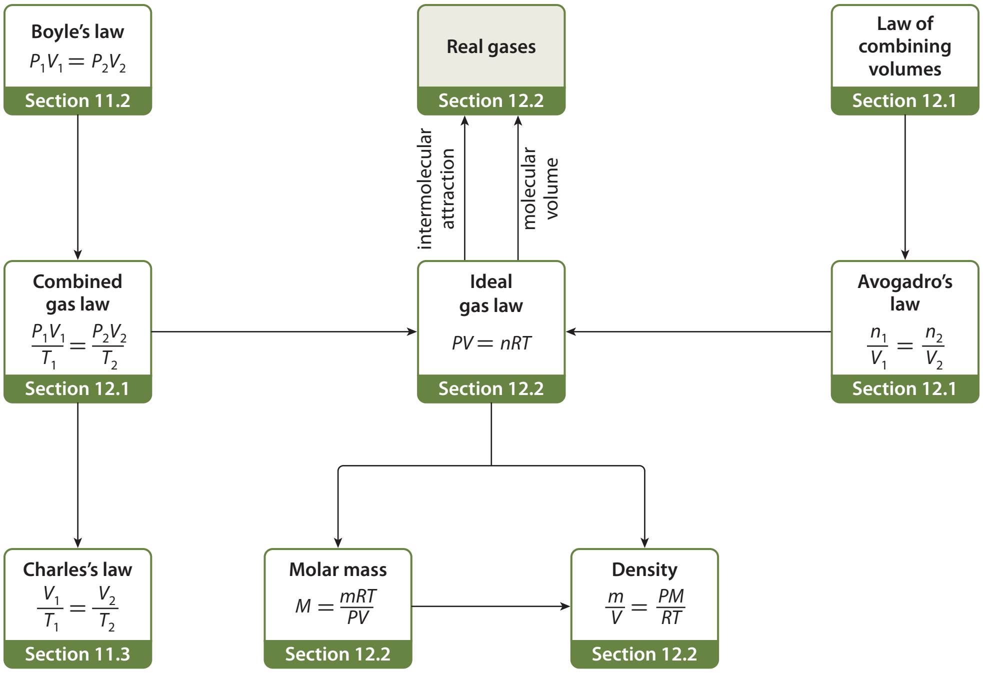
Fishbone Diagram

A fishbone diagram looks similar to a spider map, but it organizes information differently. A main topic, situation, or idea is placed in the middle of the diagram. This is the “backbone” of the “fish.” The “bones” (lines) that shoot out from the backbone can be used to list reasons why the situation exists, factors that affect the main idea, or arguments that support the main idea. Finally, supporting details shoot outward from these issues. Fishbone diagrams are useful for planning and organizing a research project. You can clearly see when you do not have enough details to support an issue, which indicates that you need to do additional research.



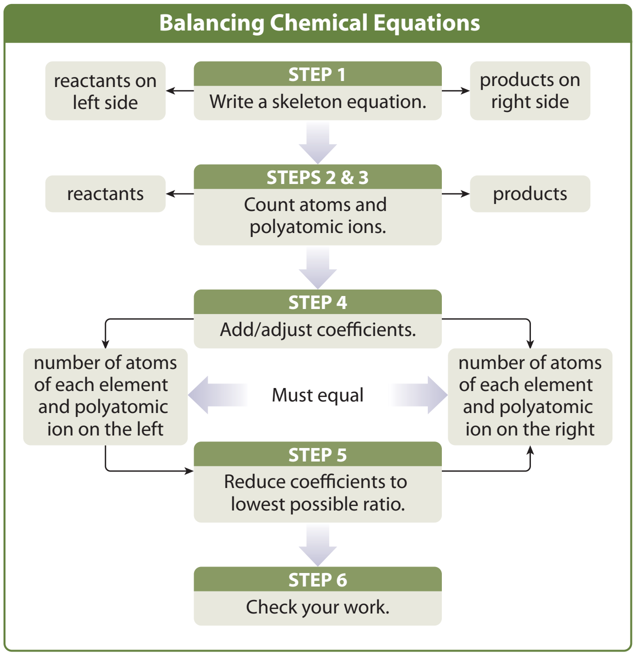
Concept Map

A concept map uses shapes and lines to show how ideas, concepts, or formulas are related. Each idea, concept, or formula is written inside a circle, a square, a rectangle, or another shape. Lines and arrows that connect the shapes indicate the relationships between them. In some cases, words that explain how the concepts are related are written on the lines that connect the shapes.



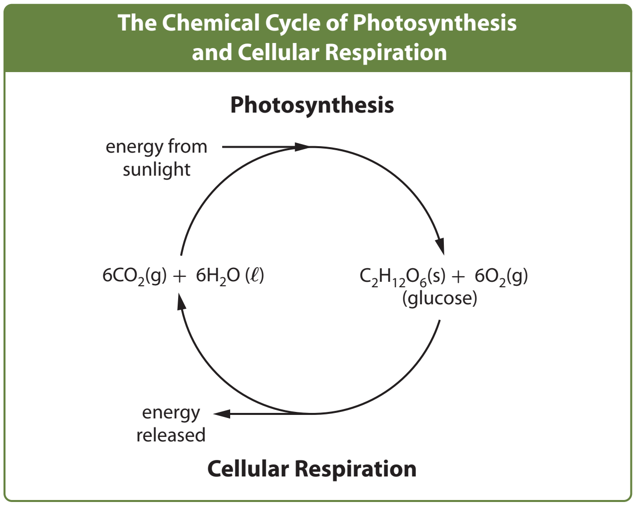
Flowchart

A flowchart shows a sequence of events or the steps in a process. An arrow leads from an initial event or step to the next event or step, and so on, until the final outcome is reached. Side arrows may also be added to provide further explanation. All the events or steps are shown in the order in which they occur.


Balancing Chemical Equations

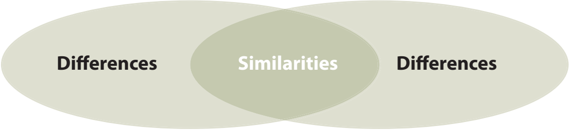
Cycle Chart

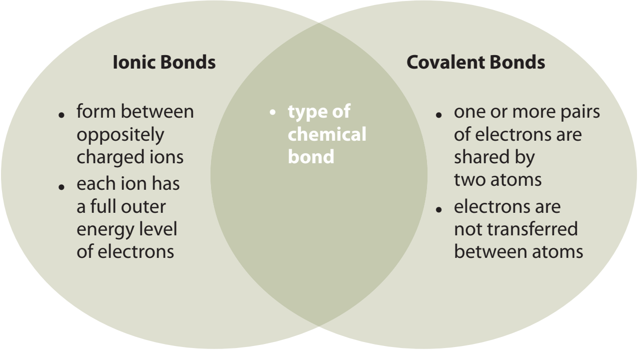
A cycle chart is a flowchart that has no distinct beginning or end. All the events are shown in the order in which they occur, as indicated by arrows, but there are no first and last events. Instead, the events occur again and again in a continuous cycle. In the photosynthesis/cellular respiration cycle, shown below, arrows branch off to show energy entering and leaving the cycle.


The Chemical Cycle of Photosynthesis and Cellular Respiration

Venn Diagram

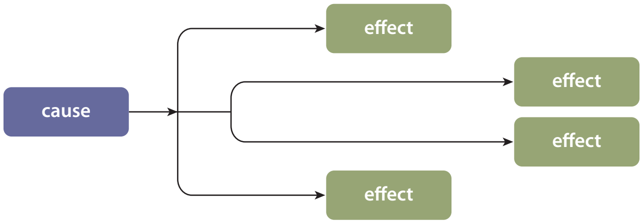
A Venn diagram uses overlapping shapes to show similarities and differences among concepts.

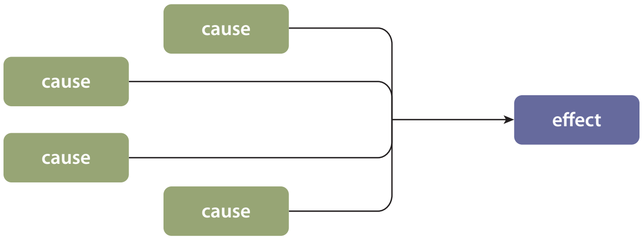




Cause-and-Effect Map

The first cause-and-effect map below shows one cause that results in several effects. The second map shows one effect that has several causes.





Instant Practice

1. Find an example of a flowchart in this textbook and discuss with a partner its effectiveness in communicating information.
2. Create a Venn diagram to compare and contrast polar and non-polar molecules.

Appendix A

Measurement

Scientists have developed globally agreed-upon standards for measurement, and for recording and calculating data. These are the standards that you will use throughout this science program.

Units of Measurement

When you take measurements for scientific purposes, you use the International System of Measurement (commonly know as SI, from the French *Système international d'unités*). SI includes the metric system and other standard units, symbols, and prefixes, which are reviewed in the tables on this page.

In SI, the base units include the metre, the kilogram, and the second. The size of any particular unit can be determined by the prefix used with the base unit. Larger and smaller units of measurement can be obtained by either multiplying or dividing the base unit by a multiple of 10.

For example, the prefix *kilo-* means multiplied by 1000. So, one kilogram is equivalent to 1000 grams:

$$1\text{ kg} = 1000\text{ g}$$

The prefix *milli-* means divided by 1000. So, one milligram is equivalent to one thousandth of a gram:

$$1\text{ mg} = \frac{1}{1000\text{ g}}$$

The following tables show the most commonly used metric prefixes, as well as some common metric quantities, units, and symbols.

Commonly Used Metric Prefixes

Prefix	Symbol	Relationship to the Base Unit
tera-	T	$10^{12} = 1\ 000\ 000\ 000\ 000$
giga-	G	$10^{9} = 1\ 000\ 000\ 000$
mega-	M	$10^{6} = 1\ 000\ 000$
kilo-	k	$10^{3} = 1\ 000$
hecto-	h	$10^{2} = 100$
deca-	da	$10^{1} = 10$
—	—	$10^{0} = 1$
deci-	d	$10^{-1} = 0.1$
centi-	c	$10^{-2} = 0.01$
milli-	m	$10^{-3} = 0.001$
micro-	µ	$10^{-6} = 0.000\ 001$
nano-	n	$10^{-9} = 0.000\ 000\ 001$
pico-	p	$10^{-12} = 0.000\ 000\ 000\ 001$

Commonly Used Metric Quantities, Units, and Symbols

Quantity	Unit	Symbol
Length	nanometre micrometre millimetre centimetre metre kilometre	nm µm mm cm m km
Mass	gram kilogram tonne	g kg t
Area	square metre square centimetre hectare	m^2 cm^2 ha (10 000 m^2)
Volume	cubic centimetre cubic metre millilitre litre	cm^3 m^3 mL L
Time	second	s
Temperature	degree Celsius	°C
Force	N	newton
Energy	joule kilojoule*	J kJ
Pressure	pascal kilopascal**	Pa kPa
Electric current	ampere	A
Quantity of electric charge	coulomb	C
Frequency	hertz	Hz
Power	watt	W

* Many dieticians in North America continue to measure nutritional energy in Calories, also known as kilocalories or dietetic Calories. In SI units, 1 Calorie = 4.186 kJ.

** In current North American medical practice, blood pressure is measured in millimetres of mercury, symbolized as mmHg. In SI units, 1 mmHg = 0.133 kPa.

Accuracy and Precision

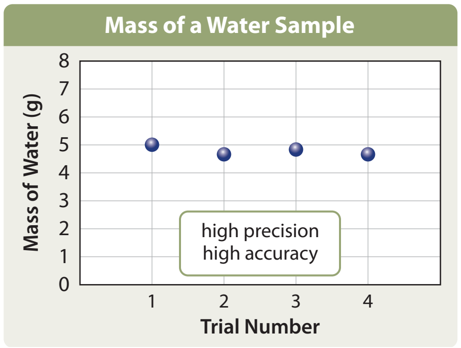
In science, the terms accuracy and precision have specific definitions that differ from their everyday meanings.

Scientific *accuracy* refers to how close a given quantity is to an accepted or expected value. For example, under standard (defined) conditions of temperature and pressure, 5 mL of water has a mass of 5 g. When you measure the mass of 5 mL of water under the same conditions, you should, if you are accurate, find the mass is 5 g.

Scientific *precision* refers to the exactness of your measurements. The precision of your measurements is directly related to the instruments you use to make the measurements. While faulty instruments (for example, a balance that is not working properly) will likely affect both the accuracy and the precision of your measurements, the calibration of the instruments you use is the factor that most affects precision. For example, a ruler calibrated in millimetres will allow you to make more precise measurements than one that shows only centimetres.

Precision also describes the repeatability of measurements. The closeness of a series of data points on a graph is an indicator of repeatability. Data that are close to one another, as in graph A, below, are said to be precise.

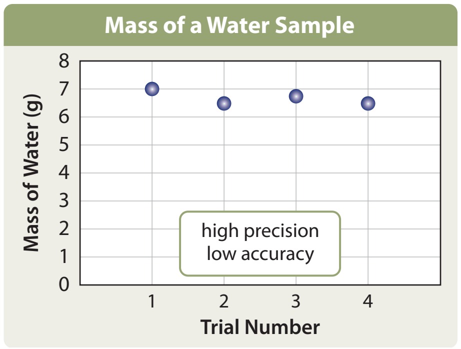
A



Graph A shows a group of data with high accuracy, since the data points are all grouped around 5 g.

There is no guarantee, however, that the data are accurate until a comparison with an accepted value is made. For example, graph B shows a group of measurements that are precise, but not accurate, since they report the mass of a 5 g sample of water as approximately 7 g.

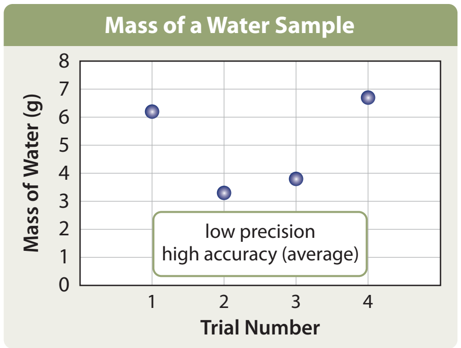
B



Graph B shows data with low accuracy, since the data points are grouped around 7 g.

In graph C, the data points give an accurate value for average mass, but they are not precise.

C



In graph C, the individual data points are not very accurate, since they are all more than 1 g away from the expected value of 5 g. However, taken as a group, the data set in graph C has high accuracy, since the average mass from the four trials is 5 g.

Error

Error exists in every measured or experimentally obtained value. Even the most careful scientist cannot avoid having error in a measurement. *Random error* results from uncontrollable variation in how we obtain a measurement. For example, human reflexes vary, so it is not possible to push the stem of a stopwatch exactly the same way every time. No measurement is perfect. Repeating trials will reduce but never eliminate the effects of random error. Random error affects precision and, usually, accuracy.

Systematic error results from consistent bias in observation. For example, a scale might consistently give a reading that is 0.5 g heavier than the actual mass of a sample, or a person might consistently read the scale of a measuring instrument incorrectly. Repeating trials will not reduce systematic error. Systematic error affects accuracy.

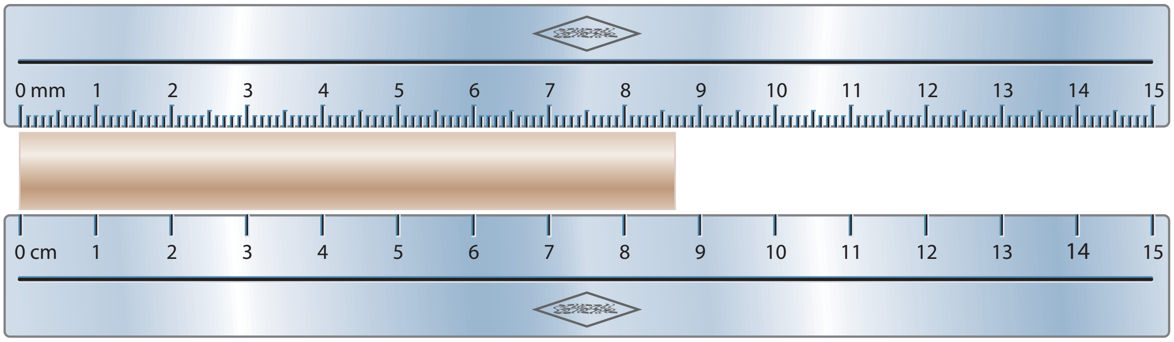
Percent Error

The amount of error associated with a measurement can be expressed as a percentage, which can help you to evaluate the accuracy of your measurement. The higher the *percent error* is, the less accurate the measurement. Percent error is calculated using the following equation:

$$\text{Percent error} = \left| \frac{\text{measured} - \text{expected value}}{\text{expected value}} \right| \times 100\%$$

(Note that the vertical lines surrounding the fraction mean *the absolute value of* the expression within the lines. That is, the expression's numerical value should be reported without a positive or negative sign.) As an example, a student measures a 5 mL sample of water and finds the mass to be 4.6 mL.

$$\begin{aligned}\text{percent error} &= \left| \frac{4.6\text{ mL} - 5\text{ mL}}{5\text{ mL}} \right| \times 100\% \\ &= \left| \frac{-0.4\text{ mL}}{5\text{ mL}} \right| \times 100\% \\ &= 8\%\end{aligned}$$



Estimated uncertainty is half of the smallest visible division. In this case, the estimated uncertainty is ±0.5 mm for the top ruler and ±0.5 cm for the bottom ruler.

Uncertainty

Estimated uncertainty describes the limitations of a measuring device. It is defined as half of the smallest division of the measuring device. For example, a metre stick with only centimetres marked on it would have an error of ± 0.5 cm. A ruler that includes millimetre divisions would have a smaller error of ± 0.5 mm (0.05 cm, or a 10-fold decrease in error). A measurement can be recorded with its estimated uncertainty. In the diagram at the top of the page, for example, the top ruler gives a measurement of 8.69 ± 0.05 cm, while the bottom ruler gives a measurement of 8.7 ± 0.5 cm.

You can convert the estimated uncertainty into a percentage of the actual measured value using the following equation:

$$\text{Relative uncertainty} = \frac{\text{estimated uncertainty}}{\text{actual measurement}} \times 100\%$$

Example

Convert the error represented by 22.0 ± 0.5 cm to a percentage.

$$\begin{aligned}\text{Relative uncertainty} &= \frac{0.5 \text{ cm}}{22.0 \text{ cm}} \times 100\% \\ &= 2\%\end{aligned}$$

Estimating

Sometimes it is not practical or possible to make an accurate measurement of a quantity. You must instead make an *estimate*—an informed judgement that approximates a quantity. For example, if you were conducting an experiment to compare the number of weeds in a field treated with herbicide with the number of weeds in an untreated field, counting the weeds would be impractical, if not impossible. Instead, you could count the number of weeds in a typical square metre of each field. You could then estimate the number of weeds in the entire field by multiplying the number of weeds in a typical square metre by the number of square metres in the field. To make a reasonable estimate of the number of weeds in the field, though, you would need to sample many areas, each 1 m^2, and then calculate an average to determine the number of weeds in a typical square metre for each field.

Estimating can be a valuable tool in science. It is important to keep in mind, however, that the number of samples you take can greatly influence the reliability of your estimate. To make a good estimate, include as many samples as is practical.

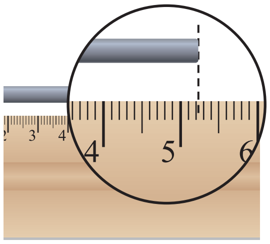
Instant Practice

1. Your teacher gives you a 500 g sample of sugar and asks you to measure its mass on a balance. You measure the sample three times, producing the measurements 492.8 g, 503.1 g, and 505.4 g. Analyze these results in terms of accuracy and precision.
2. Calculate the percent error for each of the measurements in question 1, and for the average of the three measurements. How does the accuracy of the individual measurements differ from the accuracy of the group of measurements?
3. The estimated uncertainty of the measurements in question 1 is ± 0.5 cm. Calculate the relative uncertainty of the average you determined in question 2.

Appendix A

Significant Digits and Rounding

You might think that a measurement is an exact quantity. In fact, all measurements involve uncertainty. The measuring device is one source of uncertainty, and you, as the reader of the device, are another. Every time you take a measurement, you are making an estimate by interpreting the reading. For example, the illustration below shows a ruler measuring the length of a rod. The ruler can give quite an accurate reading, since it is divided into millimetre marks. But the end of the rod falls between two marks. There is still uncertainty in the measurement. You can be certain that if the ruler is accurate, the length of the rod is between 5.2 cm and 5.3 cm. However, you must estimate the distance between the 2 mm and 3 mm marks.



Significant Digits

Significant digits are the digits you record when you take a measurement. The significant digits in a measured quantity include all the certain digits plus the first uncertain digit. In the example above, suppose you estimate the length of the rod to be 5.23 cm. The first two digits (5 and 2) are certain (those marks are visible), but the last digit (0.03) is estimated. The measurement 5.23 cm has three significant digits.

Determining the Number of Significant Digits

The following rules will help you determine the number of significant digits in a given measurement.

1. All non-zero digits (1–9) are significant.
   Examples:
   - 123 m—three significant digits
   - 23.56 km—four significant digits
2. Zeros between non-zero digits are also significant.
   Examples:
   - 1207 m—four significant digits
   - 120.5 km/h—four significant digits
3. Any zero that follows a non-zero digit *and* is to the right of the decimal point is significant.
   Examples:
   - 12.50 m/s^2—four significant digits
   - 6.0 km—two significant digits
4. Zeros that are to the left of a measurement are not significant.
   Examples:
   - 0.056—two significant digits
   - 0.007 60—three significant digits
5. Zeros used to indicate the position of the decimal are not significant. These zeros are sometimes called spacers.
   Examples:
   - 500 km—one significant digit (the decimal point is assumed to be after the final zero)
   - 0.325 m—three significant digits
   - 0.000 34 km—two significant digits
6. In some cases, a zero that appears to be a spacer is actually a significant digit. All counting numbers have an infinite number of significant digits.
   Examples:
   - 6 apples—infinite number of significant digits
   - 125 people—infinite number of significant digits
   - 450 deer—infinite number of significant digits

Instant Practice

Determine the number of significant digits in each measurement.

a. 32 individuals
b. 891 m
c. 15.764 g
d. 0.0280 L
e. 3690 km
f. 0.742 kg
g. 50.8 cm

Using Significant Digits in Mathematical Operations

When you use measured values in mathematical operations, the calculated answer cannot be more certain than the measurements on which it is based. Often the answer on your calculator will have to be rounded to the correct number of significant digits.

Rules for Rounding

1. When the first digit to be dropped is less than 5, the preceding digit is not changed.
   Example:
   - 6.723 m rounded to two significant digits is 6.7 m. The digit after the 7 is less than 5, so the 7 does not change.
2. When the first digit to be dropped is 5 or greater, the preceding digit is increased by one.
   Example:
   - 7.237 m rounded to three significant digits is 7.24 m. The digit after the 3 is greater than 5, so the 3 is increased by one.
3. When the first digit to be dropped is 5, and there are no following digits, increase the preceding number by 1 if it is odd, but leave the preceding number unchanged if it is even.
   Examples:
   - 8.345 L rounded to three significant digits is 8.34 L, because the digit before the 5 is even.
   - 8.375 L rounded to three significant digits is 8.38 L, because the digit before the 5 is odd.

Adding or Subtracting Measurements

Perform the mathematical operation, and then round off the answer so it has the same number of decimal places as the value that has the fewest decimal places.

Example:

Add the following measured lengths and express the answer to the correct number of significant digits.

$$\begin{aligned} x &= 2.3 \text{ cm} + 6.47 \text{ cm} + 13.689 \text{ cm} \\ &= 22.459 \text{ cm} \\ &= 22.5 \text{ cm} \end{aligned}$$

Since 2.3 cm has only one decimal place, the answer can have only one decimal place.

Multiplying or Dividing Measurements

Perform the mathematical operation, and then round off the answer so it has the same number of significant digits as the value that has the least number of significant digits.

Example:

Multiply the following measured lengths and express the answer to the correct number of significant digits.

$$\begin{aligned} x &= (2.342 \text{ m})(0.063 \text{ m})(306 \text{ m}) \\ &= 45.149\,076 \text{ m}^3 \\ &= 45 \text{ m}^3 \end{aligned}$$

Since 0.063 m has only two significant digits, the final answer must also have two significant digits.

Instant Practice

Perform the following calculations, rounding off your answer to the correct number of significant digits.

a. 9.745 km − 4.2 km

b. 8.33 L + 0.4 L + 56.358 L

c. 16.9 g × 0.007 56 g

d. 463.8 mL/0.660 mL

e. 580.62 mm × 1.02 mm/0.7 mm

Appendix A

Scientific Notation

An exponent is the symbol or number denoting the power to which another number or symbol is to be raised. The exponent shows the number of repeated multiplications of the base. In 10^2, the exponent is 2 and the base is 10. The expression 10^2 means 10×10.

Powers of 10

Digits	Standard Form	Exponential Form
Ten thousands	10 000	10^4
Thousands	1 000	10^3
Hundreds	100	10^2
Tens	10	10^1
Ones	1	10^0
Tenths	0.1	10^{-1}
Hundredths	0.01	10^{-2}
Thousandths	0.001	10^{-3}
Ten thousandths	0.0001	10^{-4}

Why use exponents? Consider this: One molecule of water has a mass of 0.000 000 000 000 000 000 000 029 9 g. Using such a number for calculations would be quite awkward. The mistaken addition or omission of a single zero would make the number either 10 times larger or 10 times smaller than it actually is. Scientific notation allows scientists to express very large and very small numbers more easily, to avoid mistakes, and to clarify the number of significant digits.

Expressing Numbers in Scientific Notation

In scientific notation, a number has the form $x \times 10^n$, where x is greater than or equal to 1 but less than 10, and 10^n is a power of 10. To express a number in scientific notation, use the following steps:

1. To determine the value of x, move the decimal point in the number so that only one non-zero digit is to the left of the decimal point.
2. To determine the value of the exponent n, count the number of places the decimal point moves to the left or right. If the decimal point moves to the right, express n as a positive exponent. If the decimal point moves to the left, express n as a negative exponent.
3. Use the values you have determined for x and n to express the number in the form $x \times 10^n$.

Examples

Express 0.000 000 000 000 000 000 000 029 9 g in scientific notation.

1. To determine x, move the decimal point so that only one non-zero number is to the left of the decimal point:

   2.99

2. To determine n, count the number of places the decimal moved:

   0.000 000 000 000 000 000 000 02.99 g

   3 6 9 12 15 18 21 23

   Since the decimal point moved to the right, the exponent will be negative.

3. Express the number in the form $x \times 10^n$:

   2.99×10^{-23} g

Express 602 000 000 000 000 000 000 000 in scientific notation.

1. To determine x, move the decimal point so that only one non-zero number is to the left of the decimal point:

   6.02

2. To determine n, count the number of places the decimal moved:

   6.02 000 000 000 000 000 000 000.

   23 21 18 15 12 9 6 3

   Since the decimal point moved to the left, the exponent will be positive.

3. Express the number in the form $x \times 10^n$:

   6.02×10^{23}

Logarithms and Calculating pH

An understanding of logarithms is essential for calculating the pH of a solution.

Logarithms

The logarithm of a number is the power to which you must raise a base to equal that number. By convention, we usually use 10 as the base. Every positive number has a logarithm. For example, the logarithm of 10 is 1, because $10^1 = 10$. The logarithm of 100 is 2, because $10^2 = 100$. This can be understood by examining the following equation:

$$\log_a x = y; \text{ where } a^y = x$$

Therefore, since $10^1 = 10$, we know $\log_{10} 10 = 1$. This can also be written as $\log 10 = 1$, since it is understood that 10 is used as the base by convention unless otherwise indicated. Similarly, since $10^2 = 100$, we know $\log_{10} 100 = 2$ or $\log 100 = 2$.

All numbers that are greater than 1 have a positive logarithm. Numbers that are between 0 and 1 have a negative logarithm. For instance, since $10^{-3} = 0.001$, we know $\log 0.001 = -3$. Note also that the number 1 has a logarithm of 0. The table below shows several examples of numbers and their logarithms.

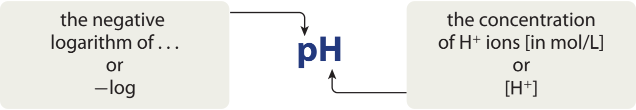
Some Numbers and Their Logarithms

Number	Scientific Notation	As a Power of 10	Logarithm
1 000 000	1×10^{6}	10^{6}	6
7 895 900	7.8959×10^{6}	$10^{6.897\,40}$	6.897 40
1	1×10^{0}	10^{0}	1
0.000 001	1×10^{-5}	10^{-5}	−5
0.004 276	4.276×10^{-3}	$10^{-2.3690}$	−2.3690

Logarithms are a convenient method for communicating large and small numbers, and are especially useful for expressing values that span a range of powers of 10. For instance, the Richter scale for earthquakes, the decibel scale for sound, and the pH scale for acids and bases all use logarithmic scales.

Calculating pH

The pH of an acid solution is defined as $-\log[H^+]$, where the square brackets mean *concentration*. In the figure below, the "p" in pH represents "the negative logarithm of..." As the logarithm of a number refers to an exponent or "power," the "p" can be thought of as "power." The power referred to is exponential power: the power of 10. Similarly, the "H" stands for the concentration of hydrogen ions, measured in mol/L.



The concept of pH allows hydrogen ion concentrations to be expressed as positive numbers, rather than negative exponents. For example, the $[H^+]$ of neutral water at 25°C is 1.0×10^{-7}. You can find the pH of water (and of any solution) by taking the negative log of the concentration of hydrogen ions.

$$\begin{aligned} \therefore \text{pH} &= -\log\,[H^+] \\ &= -\log\,(1.0 \times 10^{-7}) \\ &= -(-7.00) \\ &= 7.00 \end{aligned}$$

For a logarithm, only the digits to the right of the decimal place are significant. The numbers to the left of the decimal place reflect the power of base 10, and are, therefore, not significant.

Example:

1. Find the pH of a solution with a hydrogen ion concentration of 0.004 76 mol/L.

$$\begin{aligned} \text{pH} &= -\log\,[H^+] \\ &= -\log\,(0.004\,76 \text{ mol/L}) \\ &= 2.322 \end{aligned}$$

Note that the pH scale is a negative log scale. Thus, a decrease from pH 7 to pH 4 is actually an increase of 10^3, or 1000 times, in the acidity of a solution. An increase from pH 3 to pH 6 is a decrease in acidity of 10^3 times.

Instant Practice

1. Find the pH of a solution with the following hydrogen ion concentrations.
   - **a.** 0.000 000 01 mol/L
   - **b.** 8.7×10^{-3} mol/L

Appendix A

Preparing Solutions

Using a Volumetric Flask to Prepare a Standard Aqueous Solution

1. Place the known mass of solute in a clean beaker. Use distilled water to dissolve the solute completely.	
2. Rinse a clean volumetric flask of the required volume with a small quantity of distilled water. Discard the rinse water. Repeat the rinsing several times.	
3. Transfer the solution from the beaker to the volumetric flask using a funnel.	
4. Using a wash bottle, rinse the beaker with distilled water, and pour the rinse water into the volumetric flask. Repeat this rinsing several times.	
5. Using a wash bottle or a beaker, add distilled water to the volumetric flask until the level is just below the graduation mark. Then remove the funnel from the volumetric flask.	
6. View the neck of the volumetric flask straight on from the side, so that the graduation mark looks like a line, not an ellipse. Add distilled water, drop by drop, until the bottom of the *meniscus* (the curved surface of the solution) appears to touch the graduation mark.	

Using a Volumetric Pipette to Measure the Volume of a Stock Solution for Dilution

1. Make sure that the outside of the pipette, especially the tip, is dry. If not, wipe it with a paper towel.	
2. Squeeze the pipette bulb, and then place it over the top of the pipette. If using a pipette pump, place it over the top of the pipette.	
3. Rinse the pipette as follows. Place the tip of the pipette below the surface of the stock solution. Release the bulb carefully to draw up some liquid until the pipette is about half full. Remove the pipette bulb, invert the pipette, and drain the liquid into a beaker for waste. Repeat this rinsing two or three times.	
4. Fill the pipette with stock solution so that the level is past the graduation mark, but do not allow stock solution into the pipette bulb.	
5. Remove the pipette bulb, and quickly seal the top of the pipette with your finger or thumb.	
6. Remove the tip of the pipette from the stock solution. Lift your finger slightly, and let stock solution drain out slowly until the meniscus reaches the graduation mark. Wipe the tip of the pipette with a piece of paper towel.	
7. Move the pipette to the container into which you want to transfer the stock solution. Touch the tip of the pipette against the inside of the container, and release your finger to allow the liquid to drain. A small volume of liquid will remain inside the pipette. The pipette has been calibrated to allow for this volume of liquid. *Do not* force this liquid from the pipette.	

Appendix A

Performing an Acid-base Titration

The following steps describe how to prepare for and perform a titration.

Rinsing the Volumetric or Graduated Pipette

Rinse a pipette with the solution whose volume you are measuring. This will ensure that any drops remaining inside the pipette will form part of the measured volume.

Figure A Squeeze the pipette bulb as you put it on the stem of the pipette.

1. Put the pipette bulb on the pipette, as shown in **Figure A**. Place the tip of the pipette into a beaker of distilled water.
2. Relax your grip on the bulb to draw up a small volume of distilled water.
3. Remove the bulb, and discard the water by letting it drain out.
4. Pour a sample of the solution with the unknown concentration into a clean, dry beaker.
5. Rinse the pipette by drawing several millilitres of the solution with the unknown concentration from the beaker into the pipette. Coat the inner surface with the solution, as shown in **Figure B**. Discard the rinse. Rinse the pipette twice in this way. The pipette is now ready to be filled with the solution that has the unknown concentration.

Figure B Cover the ends of the pipette so that none of the solution spills out as you rock the pipette back and forth to coat its inner surface with solution.

Filling the Pipette

6. Place the tip of the pipette below the surface of the solution with the unknown concentration.
7. Hold the suction bulb loosely on the end of the glass stem. Use the suction bulb to draw the solution up to the point shown in **Figure C**.
8. As quickly and smoothly as you can, slide the bulb off the glass stem and place your index finger over the end.
9. Roll your finger slightly away from end of the stem to let the solution slowly drain out.
10. When the bottom of the meniscus aligns with the etched mark, as in **Figure D**, press your finger back over the end of the stem. This will prevent more solution from draining out.
11. Touch the tip of the pipette to the side of the beaker to remove any clinging drops. The measured volume inside the pipette is now ready to transfer to an Erlenmeyer flask.

Transferring the Solution

12. Place the tip of the pipette against the inside glass wall of the flask, as shown in **Figure E.** Let the solution drain slowly, by removing your finger from the stem.
13. After the solution drains, wait several seconds and then touch the tip to the inside wall of the flask to remove any drops on the end. Note: Do not remove the small amount of solution shown in **Figure F**.

Figure C Start with more of the unknown solution than you need. You will drain out the excess solution in the next two steps.

Figure D Always read the volume of the solution at the bottom of the meniscus.

Figure E Draining the pipette with the tip against the wall of the flask will prevent splashing.

Adding the Indicator

14. Add two or three drops of the indicator to the flask and its contents. Do not add too much indicator. Using more indicator does not make the colour change easier to see. Also, most indicators are weak acids. Too much indicator can change the amount of base needed for the neutralization. You are now ready to prepare the apparatus for the titration.

Rinsing the Burette

15. To rinse the burette, close the tap and add about 10 mL of distilled water from a wash bottle.
16. Tip the burette to one side, and roll it gently back and forth so that the water comes in contact with all the inner surfaces.
17. Hold the burette over a sink. Let the water drain out, as shown in **Figure G**. While you do this, check that the tap does not leak. Make sure that the tap turns smoothly and easily.
18. Rinse the burette twice, with 5 to 10 mL of the titrant. Remember to open the tap to rinse the lower portion of the burette. Discard the rinse solution each time.

Filling the Burette

19. Assemble a retort stand and burette clamp to hold the burette. Place a funnel in the top of the burette, and put a beaker under the burette.
20. With the tap closed, add the solution until it is above the zero mark. Remove the funnel. Carefully open the tap. Drain the solution into the beaker until the bottom of the meniscus is at or below the zero mark.
21. Touch the tip of the burette against the beaker to remove any clinging drops. Check that the part of the burette below the tap is filled with solution and contains no air bubbles. **Figure H** shows the air bubbles that you should avoid.
22. Find the initial burette reading using a meniscus reader, as shown in **Figure I**. Record the initial volume to the nearest 0.05 mL.

Titrating the Unknown Solution

23. Replace the beaker with the Erlenmeyer flask that contains the solution you want to titrate. Place a sheet of white paper under the flask to help you see the colour change.
24. Add titrant from the burette to the Erlenmeyer flask by opening the tap, as shown in **Figure J**. You may start by adding the titrant quickly, but slow down when you start to see a colour change in the solution in the flask.
25. At first, the colour change will disappear as you mix the solution in the flask. Add a small amount of titrant, and swirl thoroughly before adding any more. Stop adding titrant when the solution in the Erlenmeyer flask has a persistent colour change. If you are using phenolphthalein as an indicator, stop when the solution is a faint pink colour.
26. Use the meniscus reader to read the final volume. Record this volume, and subtract the initial volume from it to find the volume of the titrant needed to reach the end point.

Figure J Always swirl the flask as you add the titrant. If you have trouble swirling and adding titrant at the same time, use a magnetic stirrer or have your laboratory partner swirl the flask as you add the titrant.

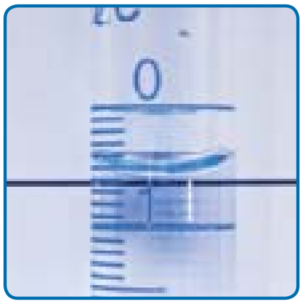


Figure I Hold the meniscus reader so that the line is under the meniscus.

Figure F A small amount of solution will always remain in the tip of the pipette. Do not remove this.

Figure G The tap is fully open when the handle on the tap is parallel to the burette and the solution inside the burette comes out quickly.

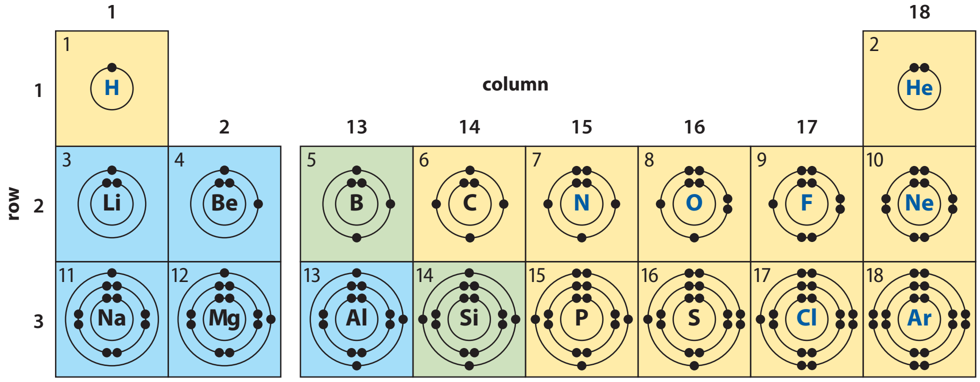
Figure H Do NOT start a titration if you have air bubbles like these in the tip of the burette. They will cause errors in your measurements.

Appendix B

Chemistry Data Tables

Electron Configurations

Bohr-Rutherford Diagrams for Elements 1-18



Names and Formulas of Ions

Ionic Charges of Representative Elements

IA 1	IIA 2	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	VIIIA 18
H^+						H^-	Noble gases do not ionize in nature.
Li	Be^{2+}			N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}	Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}					I^-	
Cs^+	Ba^{2+}						

Charges of Some Transition Metal Ions

1+	2+	3+
silver, Ag^+	cadmium, Cd^{2+} nickel, Ni^{2+} zinc, Zn^{2+}	scandium, Sc^{3+}

Common Metal Ions with More Than One Ionic Charge

Formula	Stock Name	Classical Name
Cu^+	copper(I) ion	cuprous ion
Cu^{2+}	copper(II) ion	cupric ion
Fe^{2+}	iron(II) ion	ferrous ion
Fe^{3+}	iron(III) ion	ferric ion
Hg_2^{2+} (Hg^+)	mercury(I) ion	mercurous ion
Hg^{2+}	mercury(II) ion	mercuric ion
Pb^{2+}	lead(II) ion	plumbous ion
Pb^{4+}	lead(IV) ion	plumbic ion
Sn^{2+}	tin(II) ion	stannous ion
Sn^{4+}	tin(IV) ion	stannic ion
Cr^{2+}	chromium(II) ion	chromous ion
Cr^{3+}	chromium(III) ion	chromic ion
Mn^{2+}	manganese(II) ion	
Mn^{3+}	manganese(III) ion	
Mn^{4+}	manganese(IV) ion	
Co^{2+}	cobalt(II) ion	cobaltous ion
Co^{3+}	cobalt(III) ion	cobaltic ion

Some Common Polyatomic Ions

Name	Formula
ammonium	NH_4^+
acetate or ethanoate	CH_3COO^-
benzoate	$C_6H_5COO^-$
borate	BO_3^{3-}
carbonate	CO_3^{2-}
hydrogen carbonate	HCO_3^-
perchlorate	ClO_4^-
chlorate	ClO_3^-
chlorite	ClO_2^-
hypochlorite	ClO^-
chromate	CrO_4^{2-}
dichromate	$Cr_2O_7^{2-}$
cyanide	CN^-
hydroxide	OH^-
iodate	IO_3^-
nitrate	NO_3^-
nitrite	NO_2^-
oxalate	$OOCCOO^{2-}$
hydrogen oxalate	$HOOCCOO^-$
permanganate	MnO_4^-
phosphate	PO_4^{3-}
hydrogen phosphate	HPO_4^{2-}
dihydrogen phosphate	$H_2PO_4^-$
sulfate	SO_4^{2-}
hydrogen sulfate	HSO_4^-
sulfite	SO_3^{2-}
hydrogen sulfite	HSO_3^-
cyanate	CNO^-
thiocyanate	SCN^-
thiosulfate	$S_2O_3^{2-}$

Prefixes and Suffixes for Families of Polyatomic Ions

Relative Number of Oxygen Atoms	Prefix	Suffix	Example	
Family of Four				
most	per-	-ate	ClO_4^-	perchlorate
second most	(none)	-ate	ClO_3^-	chlorate
second fewest	(none)	-ite	ClO_2^-	chlorite
fewest	hypo-	-ite	ClO^-	hypochlorite
Family of Two				
most	(none)	-ate	NO_3^-	nitrate
fewest	(none)	-ite	NO_2^-	nitrite

Names and Formulas for Compounds

Rules for Naming Binary Ionic Compounds

1. The name of the metal ion is first, followed by the name of the non-metal ion.
2. The name of the metal ion is the same as the name of the metal atom.
3. If the metal is a transition metal, it might have more than one possible charge. In these cases, a roman numeral is written in brackets after the name of the metal to indicate the magnitude of the charge.
4. The name of the non-metal ion has the same root as the name of the atom, but the suffix is changed to *-ide*.

Rules for Writing Chemical Formulas for Ionic Compounds

1. Identify the positive ion and the negative ion.
2. Find the chemical symbols for the ions, either in the periodic table or in the table of polyatomic ions. Write the symbol for the positive ion first and the symbol for the negative ion second.
3. Determine the charges of the ions. If you do not know the charges, you can find them in the periodic table.
4. Check to see if the charges differ. If the magnitudes of the charges are the same, the formula is complete. If they differ, determine the number of each ion that is needed to create a zero net charge. Write the numbers of ions needed as subscripts beside the chemical symbols, with one exception. When only one ion is needed, leave the subscript blank. A blank means one. If a polyatomic ion needs a subscript, the formula for the ion must be in brackets and the subscript must be outside the brackets.

Names of Some Common Acids without Oxygen

Pure Substance (name)	Formula H(negative ion)(aq)	Classical Name hydro(root)ic acid	IUPAC Name aqueous hydrogen (negative ion)
hydrogen fluoride	HF(aq)	hydrofluoric acid	aqueous hydrogen fluoride
hydrogen cyanide	HCN(aq)	hydrocyanic acid	aqueous hydrogen cyanide
hydrogen sulfide	H_2S(aq)	hydrosulfuric acid	aqueous hydrogen sulfide

Classical Naming System for Families of Oxoacids

		Examples	
Name of Ion	Name of Acid (dissolved in water)	Name of Ion	Name of Acid (dissolved in water)
hypo(root)ite	hypo(root)ous acid	hypochlorite, ClO^-	hypochlorous acid, HClO
(root)ite	(root)ous acid	chlorite, ClO_2^-	chlorous acid, $HClO_2$
(root)ate	(root)ic acid	chlorate, ClO_3^-	chloric acid, $HClO_3$
per(root)ate	per(root)ic acid	perchlorate, ClO_4^-	perchloric acid, $HClO_4$

Rules for Naming Binary Molecular Compounds

1. Name the element with the lower group number first. Name the element with the higher group number second.
2. The one exception to the first rule occurs when oxygen is combined with a halogen. In this situation, the halogen is named first.
3. If both elements are in the same group, name the element with the higher period number first.
4. The name of the first element is unchanged.
5. To name the second element, use the root name of the element and add the suffix *-ide*.
6. If there are two or more atoms of the first element, add a prefix to indicate the number of atoms.
7. Always add a prefix to the name of the second element to indicate the number of atoms of this element in the compound. (If the second element is oxygen, an “o” or “a” at the end of the prefix is usually omitted.)

Prefixes for Binary Molecular Compounds

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Names and Formulas for Some Common Hydrocarbons

Name	Formula
methane	$CH_4(g)$
ethane	$C_2H_6(g)$
propane	$C_3H_8(g)$
butane	$C_4H_{10}(g)$
acetylene (ethyne)	$C_2H_2(g)$
benzene	$C_6H_6(\ell)$

Ion Properties

Colours of Some Common Ions in Aqueous Solution

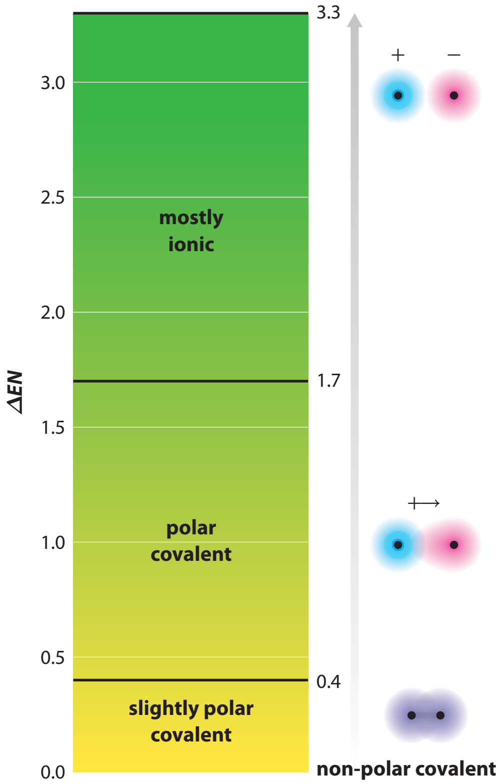
Ionic Species	Solution Concentration	
	1.0 mol/L	0.010 mol/L
chromate	yellow	pale yellow
chromium(III)	blue-green	green
chromium(II)	dark blue	pale blue
cobalt(II)	red	pink
copper(I)	blue-green	pale blue-green
copper(II)	blue	pale blue
dichromate	orange	pale orange
iron(II)	lime green	colourless
iron(III)	orange-yellow	pale yellow
manganese(II)	page pink	colourless
nickel(II)	blue-green	pale blue-green
permanganate	deep purple	purple-pink

The Flame Colour of Selected Metal Ions

Ion	Symbol	Colour
lithium	Li^+	red
sodium	Na^+	yellow
potassium	K^+	violet
cesium	Cs^+	violet
calcium	Ca^{2+}	red
strontium	Sr^{2+}	red
barium	Ba^{2+}	yellowish-green
copper	Cu^{2+}	bluish-green
baron	B^{2+}	green
lead	Pb^{2+}	bluish-white

Bond Character

Predicting Bond Character from Electronegativity Difference Values

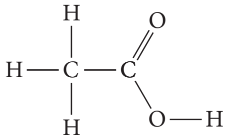


Character of Bonds

Electronegativity Difference	0.00	0.65	0.94	1.19	1.43	1.67	1.91	2.19	2.54	3.03
Percent Ionic Character	0	10	20	30	40	50	60	70	80	90
Percent Covalent Character	100	90	80	70	60	50	40	30	20	10

Types of Chemical Formulas

Empirical, Molecular, and Structural Formulas

	Empirical Formula	Molecular Formula	Structural Formula
Description	• shows the smallest whole-number molar ratio, or proportional relationship, of the elements in a compound	• shows the actual number of atoms of each element in one molecule of the compound	• shows how the atoms are connected in a chemical compound • can indicate the presence of double and triple bonds • in diagram form, bonds are represented as single, double, or triple lines joining atoms • when it is not practical to produce a diagram, a structural formula can be expressed in text only, as shown below, by listing the cluster of atoms around each carbon atom
Example:			
acetic acid or ethanoic acid	CH_2O	$C_2H_4O_2$	 or CH_3COOH

Note: The molecular formula for acetic acid is sometimes expressed as $HC_2H_3O_2$, and the structural formula is sometimes expressed as HCH_3COO, because acetic acid ionizes in aqueous solution into H^+(aq) and $C_2H_3O_2^-$(aq) or CH_3COO^-(aq). This notation is also used for other acids, such as citric acid, $C_6H_8O_7$(aq). Since citric acid is a triprotic acid, the formula can be expressed as $H_3C_6H_5O_7$(aq).

Chemical Information

Chemicals in Everyday Life

Common Name	Chemical formula and name (other names)	Physical properties	Safety concerns	Comments
acetone	$CH_3COCH_3(\ell)$ 2-propanone	clear; evaporates quickly	flammable; toxic by ingestion and inhalation	solvent; contained in some nail polish removers
acetylene	$C_2H_2(g)$ ethyne	smells sweet	highly explosive	burns very hot, with oxygen, in oxyacetylene welding torches; used to produce a wide range of synthetic products
ASA	$C_9H_8O_4(s)$ 2-acetyloxybenzoic acid acetylsalicylic acid	white crystals with a slightly bitter taste	excessive use may cause hearing loss or Reye's syndrom, especially in young people	used in Aspirin$^{\text{TM}}$ and related medicines for pain, fever, and inflammation
baking soda	$NaHCO_3(s)$ sodium hydrogen carbonate sodium bicarbonate	tiny white crystals	none	used for baking and cleaning, as an antacid and mouthwash, and in fire extinguishers
battery acid	$H_2SO_4(aq)$ sulfuric acid	clear and odourless	corrosive	used in lead-acid storage batteries (automobile batteries)
bleach	$NaOCl(aq)$ sodium hypochlorite solution	yellowish solution with a chlorine smell	toxic, strong; oxidizing agent	household chlorine bleach; used for bleaching clothes and for cleaning
bluestone	$CuSO_4 \cdot 5H_2O(s)$ copper(III) sulfate pentahydrate cupric sulfate pentahydrate	blue crystals or blue crystalline granules	toxic by ingestion; strong irritant	used in agriculture and industry, as a germicide, and for wood preservation
borax	$Na_2B_4O_7 \cdot 10H_2O(s)$ sodium borate decahydrate	white crystals	none	main source is mining; used in the glass and ceramics industries; used for making Silly Putty® and for washing clothes
carborundum	$SiC(s)$ silicon carbide	hard, black solid	none	used as an abrasive
citric acid	$C_6H_8O_7(aq)$ or $H_3C_6H_5O_7(aq)$ 3-carboxy-3-hydroxy pentanedioic acid 2-hydroxy-1,2,3-propane tricarboxylic acid	translucent crystals with a strongly acidic taste	none	used in foods and soft drinks as an acidifying agent and an antioxidant
CFCs	$CCl_2F_2(g)$, $CCl_3F(g)$, $CClF_3(g)$ chlorofluorocarbons (freons, Freon-12)	colourless, odourless gases	CFCs are now banned by the Montreal Protocol	in the past, were used as refrigerants and aerosols
charcoal/ graphite	$C(s)$ pure carbon, in a less structured form than diamond	soft grey or black solid that rubs easily onto other substances	none	used as pencil "lead" and artists' charcoal, as a de-colorizing and filtering agent, in gunpowder, and for barbeque briquettes
cream of tartar	$KHC_4H_4O_6(s)$ potassium hydrogen tartrate	white, crystalline solid	none	used as a leavening agent in baking powder
dry ice	$CO_2(s)$ solid carbon dioxide	cold white solid that sublimates	damaging to the skin and tissue after prolonged exposure	used as a refrigerant in laboratories when cold temperatures (as low as −79°C) are required

Common Name	Chemical formula and name (other names)	Physical properties	Safety concerns	Comments
Epsom salts	$MgSO_4 \cdot 7H_2O(s)$ magnesium sulfate heptahydrate	colourless cyrstals	can cause abdominal cramps and diarrhea	used as a bath salt and in cosmetics and dietary supplements; has industrial uses
ethylene	$C_2H_4(g)$ ethene	colourless gas with sweet odour and taste	flammable	used to accelerate fruit ripening and to synthesize polymers such as polystyrene; occurs naturally in plants
ethylene glycol	$C_2H_6O_2(\ell)$ or $CH_2OHCH_2OH(\ell)$ ethane-1,2-diol glycol	clear, colourless, syrupy liquid	toxic by ingestion and inhalation	used in antifreeze and cosmetics, and as a de-icing fluid for airport runways
Glauber's salt	$Na_2SO_4 \cdot 10H_2O(s)$ sodium sulfate decahydrate	large, transparent crystals, needles, or granular powder	none	a laxative; used for paper and glass making, and in solar heat storage and air conditioning; energy storage capacity more than seven times that of water
glucose	$C_6H_{12}O_6(s)$ dextrose graph sugar corn sugar	white crystals with a sweet taste	none	source of energy for most organisms
grain alcohol	$C_2H_6O(\ell)$ or $C_2H_5OH(\ell)$ ethanol ethyl alcohol	clear, volatile liquid with distinctive odour	flammable	beverage alcohol, antiseptic, laboratory/industrial solvent; produced by the fermentation of grains or fruits
gyp rock	$CaSO_4 \cdot 2H_2O(s)$ gypsum	hard, beige mineral	none	used in plaster of Paris and as a core for drywall
hydrogen peroxide	$H_2O_2(\ell)$ dihydrogen dioxide	clear, colourless liquid	damaging to skin in high concentrations	sold as 3% solution in drugstores; non-chlorine bleach often 6% H_2O_2
ibuprofen	$C_{13}H_{18}O_2(s)$ 2-[4-(2-methylpropyl)phenyl] propanoic acid p-isobutyl-hydratropic acid	white crystals	can conflict with other medications	ingredient in over-the-counter pain relievers
laughing gas	$N_2O(g)$ nitrous oxide dinitrogen oxide	colourless, mainly odourless, soluble gas	prolonged exposure causes brain damage and infertility	used as a dental anesthetic, an aerosol propellant, and to increase fuel performance in racing cars
lime	$CaO(s)$ calcium oxide quicklime hydrated lime hydraulic lime	white powder	reacts with water to produce caustic calcium hydroxide, or slaked lime, with liberation of heat	used to make cement and to clean and nullify odours in stables
limestone	$CaCO_3(s)$ calcium carbonate	soft white mineral	none	used for making lime and for building; has industrial uses
lye	$NaOH(s)$ sodium hydroxide caustic soda	white solid, found mainly in form of beads or pellets; quickly absorbs water and carbon dioxide from the air	corrosive, strong irritant	produced by the electrolysis of brine or the reaction of calcium hydroxide and sodium carbonate; has many laboratory and industrial uses; used to manufacture chemicals and make soap
malachite	$CuCO_3 \cdot Cu(OH)_2(s)$ basic copper(II) carbonate	clear, hard, bright green mineral	none	ornamental and gem stone; copper found in the ore

Common Name	Chemical formula and name (other names)	Physical properties	Safety concerns	Comments
milk of magnesia	$Mg(OH)_2(aq)$ magnesium hydroxide magnesia magma	aqueous solution of magnesium hydroxide; $Mg(OH)_2(s)$ is a white powder	harmless if used in small amounts	antacid, laxative
moth balls	$C_{10}H_8(s)$ naphthalene	white, volatile solid with an unpleasant odour	toxic by ingestion and inhalation	used to repel insects in homes and gardens, and to make synthetic resins; obtained from crude oil
MSG	$C_5H_8NNaO_4(s)$ monosodium glutamate	white, crystalline powder	may cause headaches in some people	flavour enhancer for foods in concentrations of about 0.3%
muriatic acid	$HCl(aq)$ hydrochloric acid	colourless or slightly yellow aqueous solution	toxic by ingestion and inhalation; strong irritant	has many industrial and laboratory uses; used for processing food, cleaning, and pickling
natural gas	about 85% methane, $CH_4(g)$, 10% ethane, $C_2H_6(g)$, and some propane, $C_3H_8(g)$, butane, $C_4H_{10}(g)$, and pentane, $C_5H_{12}(g)$	odourless, colourless gas	flammable and explosive; a warning odour is added to household gas as a safety precaution	used for heating, energy, and cooking; about 3% is used as a feedstock for the chemical industry
oxalic acid	$HOOCCOOH(s)$ ethanedioic acid	strongly flavoured acid; white crystals	toxic by inhalation and ingestion; strong irritant in high concentrations	occurs naturally in rhubarb, wood sorrel, and spinach; used as wood and textile bleach, rust remover, and deck cleaner; has many industrial and laboratory uses
Pepto-Bismol$^{\text{TM}}$	$C_7H_5BiO_4(s)$ (active ingredient) 2-hydroxy-2H, 4H-benzo[d]1, 3-dioxa-2-bismacyclohexan-4-one bismuth subsalicylate	pink solid or solution	may cause stomach upset if taken in excess of recommended dose	relieves digestive difficulties by coating the digestive tract and reducing acidity
PCBs	$C_{12}H_{10-x}Cl_x(\ell)$ polychlorinated biphenyls: class of compounds with two benzene rings and two or more substituted chlorine atoms	colourless liquids	highly toxic, unreactive, and persistent; cause ecological damage	used as coolants in electrical transformers
potash	$K_2CO_3(s)$ potassium carbonate Traditionally, “potash” referred to potassium carbonate, but the name is now commonly used to refer to a whole family of potassium compounds, including potassium chloride and others.	white, granular, translucent powder	solutions irritating to tissue	laboratory and industrial uses; used in special glasses, in soaps, and as a dehydrating agent
PVC	$(C_2H_3Cl)_n(s)$ polyvinyl chloride	tough, white, unreactive solid	none	used extensively as a building material

Common Name	Chemical formula and name (other names)	Physical properties	Safety concerns	Comments
road salt	$CaCl_2(s)$ calcium chloride Other salts are also commonly used as de-icing agents on roads. For example, magnesium chloride, $MgCl_2(s)$ is sometimes used in combination with calcium chloride; sodium chloride, $NaCl(s)$, is also commonly used.	white crystalline compound	none	by-product of an industrial process that produces sodium carbonate, $Na_2CO_3(s)$, from salt brine, $NaCl(aq)$, and limestone, $CaCO_3(s)$
rotten-egg gas	$H_2S(g)$ hydrogen sulfide	colourless gas with an offensive odour	highly flammable therefore high fire risk; explosive; toxic by inhalation; strong irritant to eyes and mucous membranes	obtained from sour gas (natural gas with higher than average levels of hydrogen sulfide) during natural gas production
rubbing alcohol	$(CH_3)_2CHOH(\ell)$ propan-2-ol isopropanol isopropyl alcohol	colourless liquid with a pleasant odour	flammable, therefore high fire risk; explosive; toxic by inhalation and ingestion	has industrial and medical uses
salicylic acid	$C_7H_6O_3(s)$ or $HOC_6H_4COOH(s)$ 2-hydroxybenzoic acid	white crystalline solid	damages skin in high concentrations	can be used in different amounts in foods and dyes, and in wart treatment
sand	$SiO_2(s)$ silica	large, glassy cubic crystals	toxic by inhalation; chronic exposure to dust may cause silicosis	occurs widely in nature as sand, quartz, flint, and diatomite
slaked lime	$Ca(OH)_2(s)$ calcium hydroxide	white powder that is insoluble in water	none	used to neutralize acidity in soils to make whitewash, bleaching powder, and glass
soda ash	$Na_2CO_3(s)$ sodium carbonate	white powdery crystals	none	used to manufacture glass, soaps, and detergents
sugar	$C_{12}H_{22}O_{11}(s)$ sucrose cane sugar beet sugar	cubic white crystals	none	used in foods as a sweetener; source of metabolic energy
table salt	$NaCl(s)$ sodium chloride rock salt halite	cubic white crystals	none	produced by the evaporation of natural brines and by the solar evaporation of sea water; also mined from underground sources; used in foods and for de-icing roads
Tylenol^{TM}	$C_8H_9NO_2(s)$ N-(4-hydroxyphenyl)acetamide N-acetyl-p-aminophenol acetaminophen paracetamol	colourless, slightly bitter crystals	can be toxic if an overdose is taken	pain reliever (analgesic)
TSP	$Na_3PO_4(s)$ trisodium phosphate sodium phosphate sodium orthophosphate	white crystals	toxic by ingestion; irritant to tissue; pH of 1% solution is 11.8 to 12	used as a water softener and cleaner (for example, to clean metals and to clean walls before painting); has many industrial uses

Common Name	Chemical formula and name (other names)	Physical properties	Safety concerns	Comments
vinegar	5% solution of acetic acid, CH_3COOH or $C_2H_4O_2(aq)$	clear solution with a distinctive smell	none	used for cooking and household cleaning
vitamin C	$C_6H_8O_6(s)$ L-threo-hex-2-enomo 1, 4-lactone ascorbic acid	white crystals or powder with a tart, acidic taste	none	required in diet to prevent scurvy; found in citrus fruits, tomatoes, potatoes, and green leafy vegetables
washing soda	$Na_2CO_3 \cdot H_2O(s)$ sodium carbonate monohydrate soda ash	white powdery crystals	may be irritating to skin	used for cleaning and photography, and as a food additive; has many industrial and laboratory uses
wood alcohol	$CH_3OH(\ell)$ methanol methyl alcohol	clear, colourless liquid with faint alcoholic odour	flammable; toxic by ingestion, skin absorption, and inhalation; causes blindness and death	has many industrial and household uses; used in gasoline antifreeze and as a thinner for shellac and paint; can be mixed with vegetable oil and lye to make diesel fuel

Reactivity and Solubility

Activity Series of Metals

Metal	Displaces Hydrogen…	Reactivity
lithium		most reactive
potassium		
barium		
calcium		
sodium	from cold water	
magnesium		
aluminum		
zinc		
chromium		
iron		
cadmium		
cobalt		
nickel		
tin		
lead	from acids	
hydrogen		
copper		
mercury		
silver		
platinum		
gold		least reactive

Activity Series of Halogens

Halogen	Reactivity
fluorine	most reactive
chlorine	
bromine	
iodine	least reactive

Solubility of Common Ionic Compounds in Water

	Anion +	Cation →	Solubility of Compound*
1.	Most	Alkali metal ions: Li^+, K^+, Rb^+, Cs^+, Fr^+	Soluble
	Most	hydrogen ion, H^+	Soluble
	Most	ammonium ion, NH_4^+	Soluble
2.	nitrate, NO_3^-	Most	Soluble
	acetate (ethanoate), CH_3COO^-	Ag^+	Low solubility
		Most others	Soluble
3.	chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+, Tl^+	Low solubility
		All others	Soluble
4.	fluoride, F^-	Mg^{2+}, Ca^{2+}, Ba^{2+}, Pb^{2+}	Low solubility
		Most others	Soluble
5.	sulfate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}	Low solubility
		All others	Soluble
6.	sulfide, S^{2-}	Alkali ions and H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
		All others	Low solubility
7.	hydroxide, OH^-	Alkali ions and H^+, NH_4^+, Sr^{2+}, Ba^{2+}, Tl^+	Soluble
		All others	Low solubility
8.	phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulfite, SO_3^{2-}	Alkali ions and H^+, NH_4^+	Soluble
		All others	Low solubility

*Compounds listed as soluble have solubilities of at least 1 g/100 mL of water at 25°C and 100 kPa.

Concentration Calculations

Measures of Concentration

Type of Concentration	Formula	Common Application
Concentration as a Percent • mass/volume percent	$\text{percent (m/v)} = \frac{\text{mass of solute [in grams]}}{\text{volume of solution [in millilitres]}} \times 100\%$	• intravenous solutions, such as a saline drip
• mass percent	$\text{percent (m/m)} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$	• concentration of metals in an alloy
• volume percent	$\text{percent (v/v)} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$	• solutions prepared by mixing liquids
Very Small Concentrations • parts per million	$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$	• safety limits for contaminants, such as mercury or lead in food or water
• parts per billion	$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9$	
Molar Concentration	$\text{molar concentration} = \frac{\text{amount of solute [in moles]}}{\text{volume of solution [in litres]}}$ $c = \frac{n}{V}$	• solutions used as reactants

Acids, Bases, and Indicators

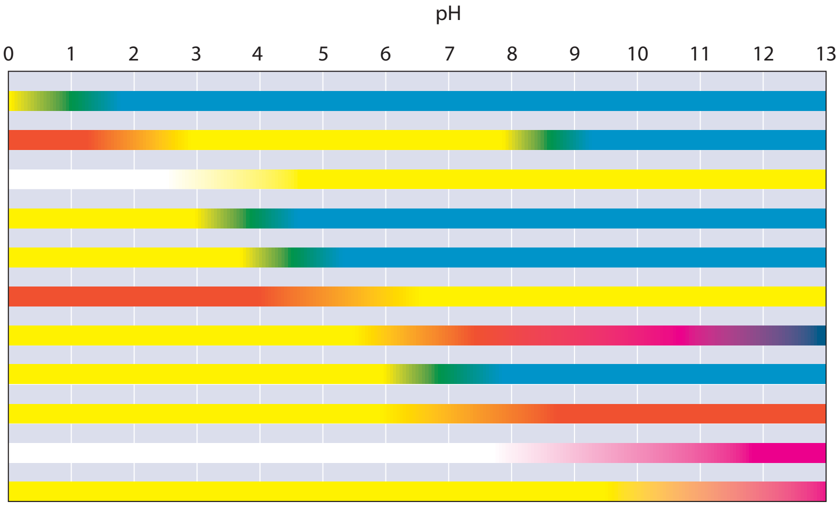
The Most Common Strong Acids

Name	Formula
hydrochloric acid	$HCl(aq)$
hydrobromic acid	$HBr(aq)$
hydroiodic acid	$HI(aq)$
perchloric acid	$HClO_4(aq)$
nitric acid	$HNO_3(aq)$
sulfuric acid	$H_2SO_4(aq)$

Some Common Strong Bases

Name	Formula
lithium hydroxide	$LiOH(aq)$
sodium hydroxide	$NaOH(aq)$
potassium hydroxide	$KOH(aq)$
calcium hydroxide	$Ca(OH)_2(aq)$
barium hydroxide	$Ba(OH)_2(aq)$

Range of Some Common pH Indicators



Indicator	pH Range	Change of Colour with Increasing pH
Crystal violet	0.0–1.6	yellow to blue
Thymol blue	1.2–2.8	red to yellow
2,4-Dinitrophenol	2.4–4.0	colourless to yellow
Bromophenol blue	3.0–4.6	yellow to blue
Bromocresol green	3.8–5.4	yellow to blue
Methyl red	4.8–6.0	red to yellow
Alizarin	5.7–7.3; 11.0–12.4	yellow to red; red to violet
Bromothymol blue	6.0–7.6	yellow to blue
Phenol red	6.6–8.0	yellow to red
Phenolphthalein	8.2–10.0	colourless to pink
Alizarin yellow R	10.1–12.0	yellow to red

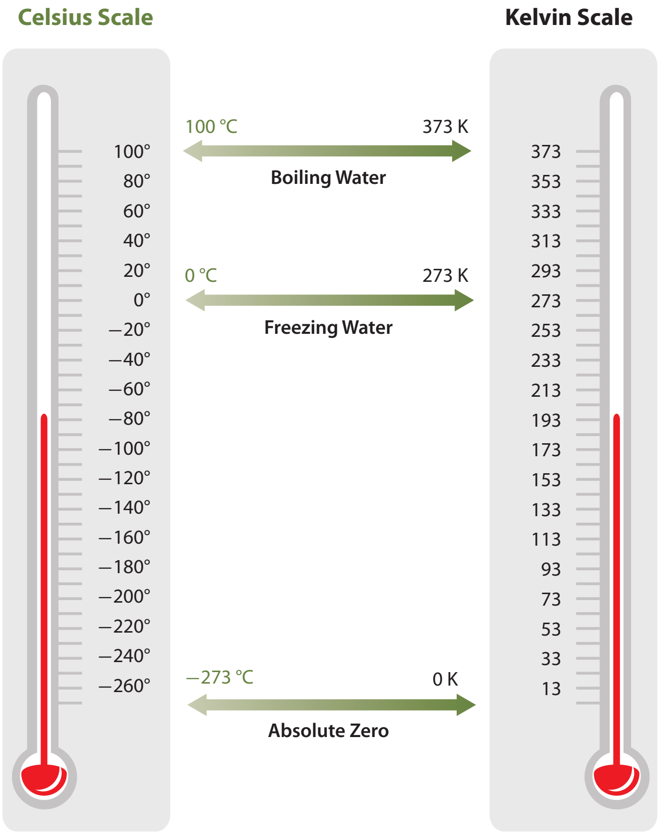
Pressure and Temperature Units

Units of Pressure Used for Various Instruments

Unit of Pressure	Symbol
standard atmosphere	atm
millimetres of mercury	mmHg
torr	torr
pascal	Pa
kilopascal	kPa
bar	bar
millibar	mb
pounds per square inch	psi

1 atm = 760 mmHg = 760 torr = 101 325 Pa = 101.325 kPa = 1.01325 bar = 14.7 psi

The Relationship between the Celsius and Kelvin Temperature Scales



For converting Celsius to kelvin: K = °C + 273.15
For converting kelvin to Celsius: °C = K − 273.15

Gases: Pressure, Volume, Temperature, and Stoichiometry Calculations

Standard Conditions of Temperature and Pressure

Conditions	Pressure	Celsius Temperature	Kelvin Temperature	Molar Volume of an Ideal Gas
STP (standard temperature and pressure)	101.325 kPa	0°C	273.15 K	22.4 L/mol
SATP (standard ambient temperature and pressure)	100.0 kPa	25°C	298.15 K	24.8 L/mol

Partial Pressures of Water Vapour at Different Temperatures

Temperature (°C)	Pressure (kPa)	Temperature (°C)	Pressure (kPa)
15	1.71	22	2.81
16	1.81	23	2.99
17	1.93	24	3.17
18	2.07	25	3.36
19	2.20	26	3.36
20	2.33	27	3.56
21	2.49	28	3.37

Universal Gas Constant

$$R = 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

Formulas for Calculations Involving Gases

	Formula
Molar mass	$M = \frac{m}{n}$
Density	$D = \frac{m}{V}$
Pressure	$P = \frac{F}{A}$
Boyle's Law	$P_1V_1 = P_2V_2$
Charles's Law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gay-Lussac's Law	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Combined Gas Law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Avogadro's Law	$\frac{n_1}{V_1} = \frac{n_2}{V_2}$
Molar volume	$v = \frac{V}{n}$
Ideal Gas Law	$PV = nRT$
Dalton's Law of Partial Pressures	$P_{\text{total}} = P_{\text{dry air}} + P_{\text{water vapour}}$

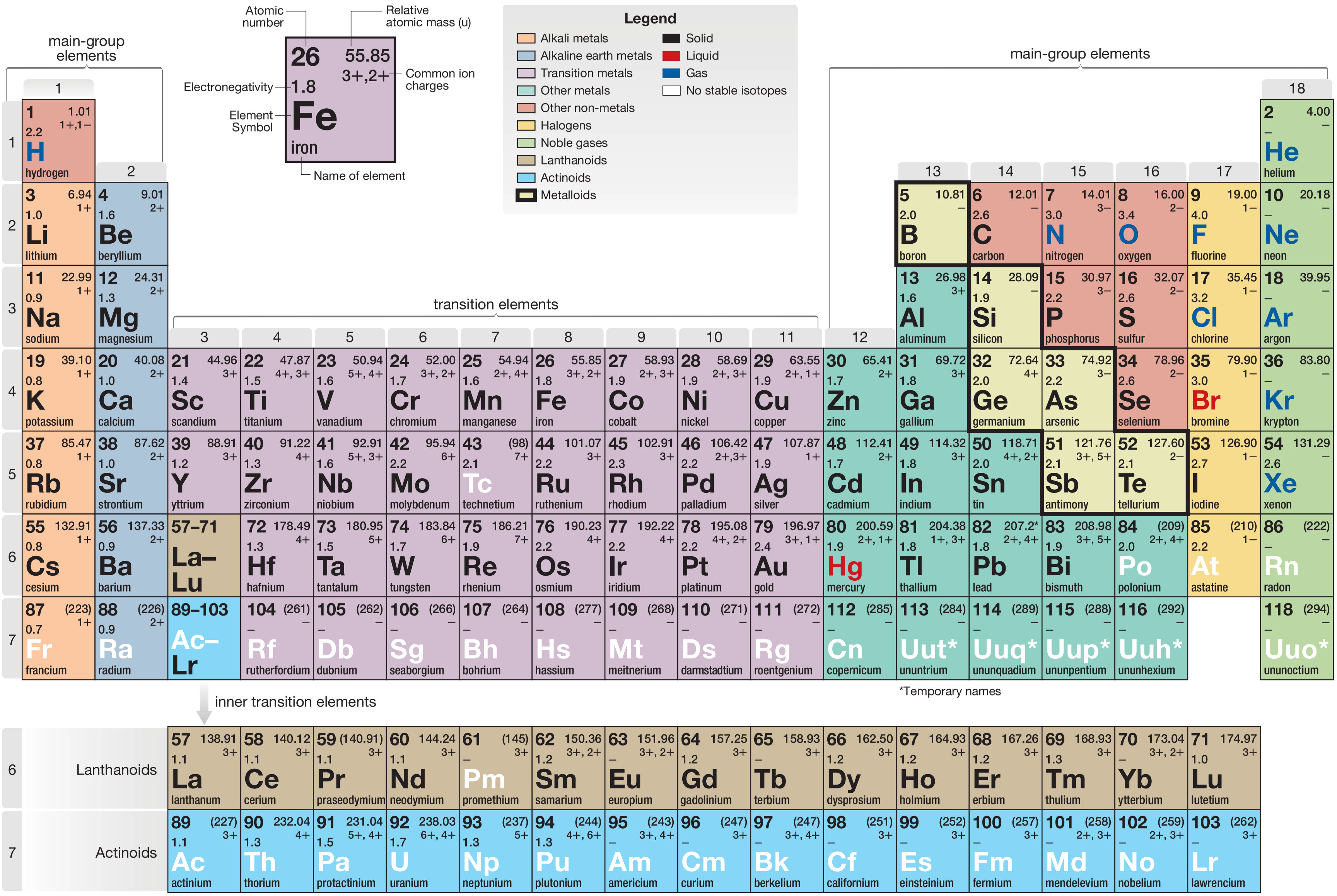
Appendix B

Alphabetical List of Elements

Element	Symbol	Atomic Number
Actinium	Ac	89
Aluminum	Al	13
Americium	Am	95
Antimony	Sb	51
Argon	Ar	18
Arsenic	As	33
Astatine	At	85
Barium	Ba	56
Berkelium	Bk	97
Beryllium	Be	4
Bismuth	Bi	83
Bohrium	Bh	107
Boron	B	5
Bromine	Br	35
Cadmium	Cd	48
Calcium	Ca	20
Californium	Cf	98
Carbon	C	6
Cerium	Ce	58
Cesium	Cs	55
Chlorine	Cl	17
Chromium	Cr	24
Cobalt	Co	27
Copernicum	Cn	112
Copper	Cu	29
Curium	Cm	96
Darmstadtium	Ds	110
Dubnium	Db	105
Dysprosium	Dy	66
Einsteinium	Es	99
Erbium	Er	68
Europium	Eu	63
Fermium	Fm	100
Fluorine	F	9
Francium	Fr	87
Gadolinium	Gd	64
Gallium	Ga	31
Germanium	Ge	32
Gold	Au	79
Hafnium	Hf	72
Hassium	Hs	108
Helium	He	2
Holmium	Ho	67
Hydrogen	H	1
Indium	In	49
Iodine	I	53
Iridium	Ir	77
Iron	Fe	26
Krypton	Kr	36
Lanthanum	La	57
Lawrencium	Lr	103
Lead	Pb	82
Lithium	Li	3
Lutetium	Lu	71
Magnesium	Mg	12
Manganese	Mn	25
Meitnerium	Mt	109
Mendelevium	Md	101
Mercury	Hg	80
Molybdenum	Mo	42
Neodymium	Nd	60
Neon	Ne	10
Neptunium	Np	93
Nickel	Ni	28
Niobium	Nb	41
Nitrogen	N	7
Nobelium	No	102
Osmium	Os	76
Oxygen	O	8
Palladium	Pd	46
Phosphorus	P	15
Platinum	Pt	78
Plutonium	Pu	94
Polonium	Po	84
Potassium	K	19
Praseodymium	Pr	59
Promethium	Pm	61
Protactinium	Pa	91
Radium	Ra	88
Radon	Rn	86
Rhenium	Re	75
Rhodium	Rh	45
Roentgenium	Rg	111
Rubidium	Rb	37
Ruthenium	Ru	44
Rutherfordium	Rf	104
Samarium	Sm	62
Scandium	Sc	21
Seaborgium	Sg	106
Selenium	Se	34
Silicon	Si	14
Silver	Ag	47
Sodium	Na	11
Strontium	Sr	38
Sulfur	S	16
Tantalum	Ta	73
Technetium	Tc	43
Tellurium	Te	52
Terbium	Tb	65
Thallium	Tl	81
Thorium	Th	90
Thulium	Tm	69
Tin	Sn	50
Titanium	Ti	22
Tungsten	W	74
Ununhexium	Uuh	116
Ununoctium	Uuo	118
Ununpentium	Uup	115
Ununquadium	Uuq	114
Ununtrium	Uut	113
Uranium	U	92
Vanadium	V	23
Xenon	Xe	54
Ytterbium	Yb	70
Yttrium	Y	39
Zinc	Zn	30
Zirconium	Zr	40

Appendix B

Periodic Table of the Elements



	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 1.01 2.2 1+,1− H hydrogen																	2 4.00 – – He helium
2	3 6.94 1.0 1+ Li lithium	4 9.01 1.6 2+ Be beryllium											5 10.81 2.0 – B boron	6 12.01 2.6 – C carbon	7 14.01 3.0 3− N nitrogen	8 16.00 3.4 2− O oxygen	9 19.00 4.0 1− F fluorine	10 20.18 – – Ne neon
3	11 22.99 0.9 1+ Na sodium	12 24.31 1.3 2+ Mg magnesium											13 26.98 1.6 3+ Al aluminum	14 28.09 1.9 – Si silicon	15 30.97 2.2 3− P phosphorus	16 32.07 2.6 2− S sulfur	17 35.45 3.2 1− Cl chlorine	18 39.95 – – Ar argon
4	19 39.10 0.8 1+ K potassium	20 40.08 1.0 2+ Ca calcium	21 44.96 1.4 3+ Sc scandium	22 47.87 1.5 4+, 3+ Ti titanium	23 50.94 1.6 5+, 4+ V vanadium	24 52.00 1.7 3+, 2+ Cr chromium	25 54.94 1.6 2+, 4+ Mn manganese	26 55.85 1.8 3+, 2+ Fe iron	27 58.93 1.9 2+, 3+ Co cobalt	28 58.69 1.9 2+, 3+ Ni nickel	29 63.55 1.9 2+, 1+ Cu copper	30 65.41 1.7 2+ Zn zinc	31 69.72 1.8 3+ Ga gallium	32 72.64 2.0 4+ Ge germanium	33 74.92 2.2 3− As arsenic	34 78.96 2.6 2− Se selenium	35 79.90 3.0 1− Br bromine	36 83.80 – – Kr krypton
5	37 85.47 0.8 1+ Rb rubidium	38 87.62 1.0 2+ Sr strontium	39 88.91 1.2 3+ Y yttrium	40 91.22 1.3 4+ Zr zirconium	41 92.91 1.6 5+, 3+ Nb niobium	42 95.94 2.2 6+ Mo molybdenum	43 (98) 2.1 7+ Tc technetium	44 101.07 2.2 3+ Ru ruthenium	45 102.91 2.3 3+ Rh rhodium	46 106.42 2.2 2+,3+ Pd palladium	47 107.87 1.9 1+ Ag silver	48 112.41 1.7 2+ Cd cadmium	49 114.32 1.8 3+ In indium	50 118.71 2.0 4+, 2+ Sn tin	51 121.76 2.1 3+, 5+ Sb antimony	52 127.60 2.1 2− Te tellurium	53 126.90 2.7 1− I iodine	54 131.29 2.6 – Xe xenon
6	55 132.91 0.8 1+ Cs cesium	56 137.33 0.9 2+ Ba barium	57–71 La–Lu	72 178.49 1.3 4+ Hf hafnium	73 180.95 1.5 5+ Ta tantalum	74 183.84 1.7 6+ W tungsten	75 186.21 1.9 7+ Re rhenium	76 190.23 2.2 4+ Os osmium	77 192.22 2.2 4+ Ir iridium	78 195.08 2.2 4+, 2+ Pt platinum	79 196.97 2.4 3+, 1+ Au gold	80 200.59 1.9 2+, 1+ Hg mercury	81 204.38 1.8 1+, 3+ Tl thallium	82 207.2* 1.8 2+, 4+ Pb lead	83 208.98 1.9 3+, 5+ Bi bismuth	84 (209) 2.0 2+, 4+ Po polonium	85 (210) 2.2 1− At astatine	86 (222) – – Rn radon
7	87 (223) 0.7 1+ Fr francium	88 (226) 0.9 2+ Ra radium	89–103 Ac–Lr	104 (261) – – Rf rutherfordium	105 (262) – – Db dubnium	106 (266) – – Sg seaborgium	107 (264) – – Bh bohrium	108 (277) – – Hs hassium	109 (268) – – Mt meitnerium	110 (271) – – Ds darmstadtium	111 (272) – – Rg roentgenium	112 (285) – – Cn copernicum	113 (284) – – Uut* ununtrium	114 (289) – – Uuq* ununquadium	115 (288) – – Uup* ununpentium	116 (292) – – Uuh* ununhexium		118 (294) – – Uuo* ununoctium

*Temporary names

6	Lanthanoids	57 138.91 1.1 3+ La lanthanum	58 140.12 1.1 3+ Ce cerium	59 (140.91) 1.1 3+ Pr praseodymium	60 144.24 1.1 3+ Nd neodymium	61 (145) – 3+ Pm promethium	62 150.36 1.2 3+, 2+ Sm samarium	63 151.96 – 3+, 2+ Eu europium	64 157.25 1.2 3+ Gd gadolinium	65 158.93 – 3+ Tb terbium	66 162.50 1.2 3+ Dy dysprosium	67 164.93 1.2 3+ Ho holmium	68 167.26 1.2 3+ Er erbium	69 168.93 1.3 3+ Tm thulium	70 173.04 – 3+, 2+ Yb ytterbium	71 174.97 1.0 3+ Lu lutetium
7	Actinoids	89 (227) 1.1 3+ Ac actinium	90 232.04 1.3 4+ Th thorium	91 231.04 1.5 5+, 4+ Pa protactinium	92 238.03 1.7 6+, 4+ U uranium	93 (237) 1.3 5+ Np neptunium	94 (244) 1.3 4+, 6+ Pu plutonium	95 (243) – 3+, 4+ Am americium	96 (247) – 3+ Cm curium	97 (247) – 3+, 4+ Bk berkelium	98 (251) – 3+ Cf californium	99 (252) – 3+ Es einsteinium	100 (257) – 3+ Fm fermium	101 (258) – 2+, 3+ Md mendelevium	102 (259) – 2+, 3+ No nobelium	103 (262) – 3+ Lr lawrencium

Although Group 12 elements are often included in the transition elements, these elements are chemically more similar to the main-group elements.
Any value in parentheses is the mass of the least unstable or best known isotope for elements that do not occur naturally.

Appendix C Answers to Selected Questions and Problems

Answers are provided for all Learning Check questions, caption questions, Practice Problems, and multiple choice questions. Answers are provided for selected Section Review, Chapter Review, Chapter Self-Assessment, Unit Review, and Unit Self-Assessment questions.

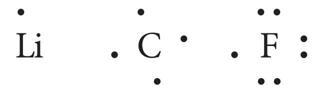
UNIT 1

Chapter 1

Answers to Learning Check Questions

1. (1) In the Thomson model of the atom, the positive charge is spread over the entire atom, whereas in the Rutherford model, the positive charge is contained in a very small volume at the centre of the atom. (2) In the Thomson model, the negative charges are embedded in the large positively charged region. In the Rutherford model, the negative charges orbit the tiny positive charge.

2.



3. The radius that Bohr calculated for the orbit of the electron in the hydrogen atom is the same as the average distance that Schrödinger calculated for the electron from the nucleus of the hydrogen atom.

4. Models represent the information that chemists have obtained about an object or concept. Chemists can use models to predict the behaviour of the object and design further experiments to test the model and, if necessary, modify the model. Also, models help chemists communicate about an object or concept.

5. Dalton: All matter consists of tiny particles (called atoms). Atoms of each element are unique. Thomson: Atoms contain negatively charged particles that can be ejected from the atom. Bohr: Electrons exist only in certain allowed energy levels in an atom.

6. $2n^2$: $2 \times 8^2 = 2 \times 64 = 128$

7. Mendeleev listed the elements vertically, in order of atomic mass (then called atomic weight). When he came to an element with properties similar to one higher in the list, he started a new column by putting the next element beside the one which had similar properties.

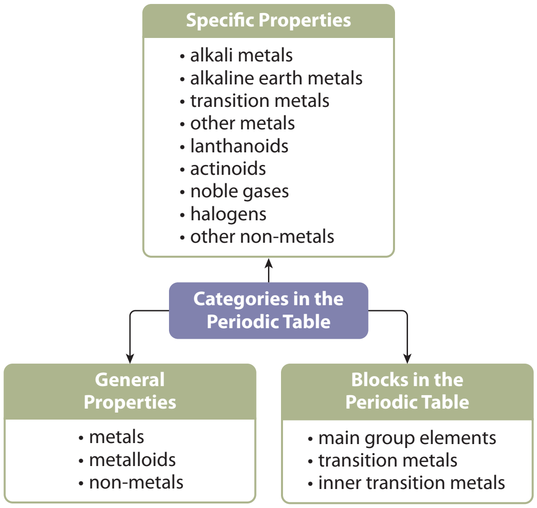
8. When elements are arranged by atomic number, their chemical and physical properties recur periodically. Many elements have similar properties and these properties follow a pattern that repeats itself regularly.

9. Each column in the periodic table constitutes a group. Groups contain elements with similar chemical and physical properties. Each row in the periodic table constitutes a period. The atomic number of the elements increases sequentially across a period The outermost electron shell that is occupied is the same for each element in a period.

10. A specific electron shell is filling as you go across a period. When the shell is filled, the period ends. Elements with filled outer electron shells are noble gases.

11. The elements in the periodic table are categorized in several different ways. In one case, elements are categorized by whether they are metals, metalloids, or non-metals. In another case, the elements are categorized by very specific chemical and physical properties. Elements are also categorized by dividing the periodic table into blocks.

12.



13. The radius of an atom is the radius of a sphere within which electrons spend 90 percent of the time.

14. Electrons exist is a region that is best described as a cloud so atoms do not have defined boundaries. There is no way to directly measure the radius within which electrons spend 90 percent of their time.

15. As the charge of a nucleus increases, it exerts a greater force on the electrons. Thus, for electrons in a given energy level, the electrons are drawn closer to the nucleus. As a result, the size of the atom decreases across a period from left to right in the periodic table.

16. Electrons in filled shells reduce the effect of positive charge on the outer electrons. Thus, outer electrons are not as strongly attracted to the nucleus as they would be if the electrons in the lower energy levels were absent. As a result, the size of an atom increases down a group in the periodic table.

17. Increasing atomic number: oxygen (8), potassium (19), krypton (36), tin (50)

Increasing size (atomic radius): oxygen (73 pm), krypton (112 pm), tin (140 pm), potassium (227 pm).

As the atomic number increases going across a period from left to right, the nuclear charge increases, which means there is more pull on the electrons and therefore the atomic radius decreases. Thus, within a period the progression of the atomic number and size are opposite. Going down a group, however, even though the atomic number increases, the effective nuclear charge is reduced due to shielding; the atomic radius therefore increases. Also, the number of occupied electron shells increases, making the atoms larger.

18. The nuclear charge; the number of occupied electron shells; shielding; the number of valence electrons

Answers to Practice Problems

1. 35.45 u
2. 10.8 u
3. 6.94 u
4. 24.31 u
5. 39.9%; 69.7 u
6. 49.31%; 79.91 u
7. 85.47 u
8. The isotopic abundance of nitrogen-14 is very high with just a trace of nitrogen-15.
9. 186.2 u
10. 193 u

Answers to Caption Questions

Figure 1.10 Columns represent periods. The length of the period depends on the number of electrons allowed in the highest energy electron shell. That number increases with period number.

Figure 1.15

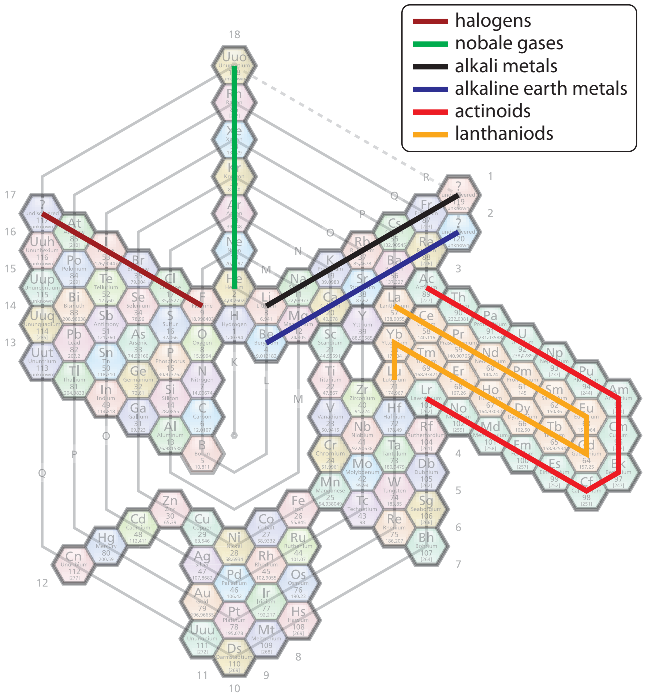


Figure 1.18 The atomic radii get smaller as you go from left to right across any period. Atomic radii get larger as you go down any group. The last three elements of each period have very similar atomic radii. The change in atomic radii across a period shows the most dramatic change between groups 2 and 13 and between groups 15 and 16.

Figure 1.22 Fluorine has the greatest electronegativity and francium has the smallest electronegativity. They are on corners of the periodic table diagonal to each other.

Figure 1.23 Nuclei that can get closer to the outer electrons of another atom will attract those electrons with a greater force. Therefore, atoms with smaller radii will have a higher electronegativity.

Answers to Section 1.1 Review Questions

9. Isotope Data

	Name of Isotope	Notation for Isotope	Atomic Number	Mass Number	Number of Protons	Number of Electrons	Number of Neutrons
a.	bromine-81	$^{81}_{35}Br$	35	81	35	35	46
b.	neon-22	$^{22}_{10}Ne$	10	22	10	10	12
c.	calcium-44	$^{44}_{20}Ca$	20	44	20	20	24
d.	silver-107	$^{107}_{47}Br$	47	107	47	47	60

10. 28.09 u

Answers to Section 1.2 Review Questions

13 a. Na **b.** Br **c.** H

Answers to Section 1.3 Review Questions

7. A^{2+} + energy $\longrightarrow A^{3+} + e^-$ (third ionization) The third ionization energy will be larger than the first or second because the electron is essentially being pulled away from a 3+ charge as opposed to 2+ and 1+ for the second and first ionization energies respectively.

14. Aluminum: 26.98 u, 1.43×10^{-10} m
Lead: 207.2 u, 1.46×10^{-10} m

15. a. 2.45×10^{-19} J

Answers to Chapter 1 Review Questions

1. c	**3.** d	**5.** b	**7.** c
2. a	**4.** b	**6.** e	**8.** b

15. A^+ + energy $\longrightarrow A^{2+} + e^-$

18. 121.76 u

Answers to Chapter 1 Self Assessment Questions

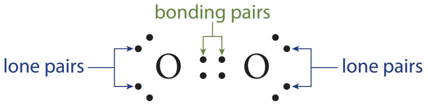
1. b	**3.** d	**5.** a	**7.** e	**9.** a
2. d	**4.** b	**6.** d	**8.** b	**10.** e

13. 152.0 μ

Chapter 2

Answers to Learning Check Questions

1. When bonds form between atoms, the atoms gain, lose, or share electrons in such a way that they create a filled outer shell containing eight electrons. For example, a fluorine atom can gain one electron and become a fluoride ion that has a completed octet of electrons in its outer shell.
2. One calcium atom can donate an electron to each of two bromine atoms. The combination of one calcium ion and two bromide ions results in a neutral compound.
3. Determine the total number of valence electrons that each of the atoms in the compound should have and add them together. Count the number of electrons shown in the Lewis structure. If the numbers are equal, the compound is neutral and is a molecular compound. If the numbers are not equal, the compound carries a charge and is thus a polyatomic ion.
4.



5. Double bonds form when four electrons are shared by two atoms. Triple bonds form when six electrons are shared by two atoms. In some cases, when two atoms share two electrons, neither atom has an octet of electrons in its outer shell. When both atoms contribute another electron, they share four electrons in total to form a double bond. They might then both have an octet of electrons in their outer shells. If not, the atoms may share another electron each to form a triple bond. Double and triple bonds are used to ensure that all the elements in the compound have an octet of electrons in their outer shell.

6. A group of two or more atoms of non-metal elements can share electrons and form covalent bonds, but, as a group, must either lose or gain electrons so that they can all form a stable octet of electrons. In this case, the atoms have formed a polyatomic ion. This ion can then form an ionic bond with other ions.

7. A binary ionic compound is an ionic compound that contains atoms of two and only two different elements.

8. **a.** K_2S, potassium sulfide
  b. MgO, magnesium oxide
  c. $FeCl_2$, iron(II) chloride, or $FeCl_3$, iron(III) chloride
  d. Mg_3N_2, magnesium nitride
  e. HI, hydrogen iodide
  f. $Ca(OH)_2$, calcium hydroxide

9. **a.** chromium(II) bromide
  b. sodium sulfide
  c. mercury(I) chloride
  d. lead(II) iodide
  e. aqueous hydrogen nitrate, or nitric acid
  f. potassium hydroxide

10. **a.** $ZnBr_2$
  b. Al_2S_3
  c. Cu_3N_2
  d. $MgCl_2$
  e. H_3N
  f. $Cu(OH)_2$

11. **a.** hypofluorite
  b. fluorite
  c. fluorate
  d. perfluorate

When there is a family of compounds that can have 1, 2, 3, or 4 oxygen atoms, the combination of prefixes and suffixes are hypo-...-ite for one oxygen atom, (none)...-ite for two oxygen atoms, (none)-...-ate for three oxygen atoms, and per-...-ate for four oxygen atoms.

12. **a.** $FeSO_4$
  b. $NaNO_3$
  c. $CuCrO_4$
  d. $Mg_3(PO_4)_2$
  e. H_2CO_3
  f. $Al(OH)_3$

13. When a substance is melting, the particles (ions or molecules) have gained enough energy to break the bonds or attractive forces between the particles.

14. The compound with a melting point of 714°C is probably an ionic compound and the one with a melting point of 146°C is probably a molecular compound.

15. Potassium iodide is an ionic compound and, in ionic compounds, each ion is attracted to every oppositely charged ion adjacent to it. There are no combinations of ions that are unique and therefore cannot be called molecules.

16. A dipole-dipole force is an attractive force between the slightly positive end of a polar molecule and the slightly negative end of another polar molecule.

17. The forces of attraction among non-polar molecules are very weak. It takes only a small amount of energy for these molecules to pull apart. This means that a relatively low temperature is capable of supplying the small amount of needed energy.

18. Intermolecular forces include dipole-dipole forces and the weak forces among non-polar molecules.

Answers to Practice Problems

1. tetraphosphorus heptasulfide
2. lead(II) nitrate
3. $MnCl_4$
4. NI_3
5. copper(I) bromide
6. Fe_2O_3
7. SiO_2
8. selenium hexafluoride
9. calcium oxide
10. $Co(NO_3)_3$

Answers to Caption Questions

Figure 2.9 When you are confident that the Lewis structure is correct and all atoms have an octet of electrons, the number of shared electrons shows the number of bonds. One pair of electrons between two atoms represents one bond.

Figure 2.17 The electronegativity of chlorine is larger than the electronegativity of carbon, indicating that the chlorine attracts the shared electrons with a greater force than does the carbon.

Figure 2.22 The common name of dihydrogen monoxide is water.

Answers to Section 2.1 Review Questions

6. 6 electrons (3 pairs of electrons)

10. **a.** $\Delta EN = 1.4$; polar covalent
  b. $\Delta EN = 0.4$; polar covalent (on the line)
  c. $\Delta EN = 0.0$; non-polar covalent
  d. $\Delta EN = 1.5$; polar covalent
  e. $\Delta EN = 0.3$; slightly polar covalent
  f. $\Delta EN = 3.1$; mostly ionic
  g. $\Delta EN = 1.6$; polar covalent
  h. $\Delta EN = 1.8$; mostly ionic

Answers to Section 2.2 Review Questions

5. **a.** aluminum oxide
  b. mercury(II) iodide
  c. sodium phosphide
  d. potassium phosphate
  e. ammonium chloride

f. lithium perchlorate
g. aqueous hydrogen nitrate or nitric acid
h. lithium hydroxide

6. **a.** ZnO
b. FeS
c. KClO
d. MgI_2
e. $CoCl_3$
f. NaCN

8. nitrogen oxide, nitrogen dioxide, dinitrogen monoxide, dinitrogen trioxide, dinitrogen tetroxide, dinitrogen pentoxide

9. **a.** PCl_5
b. F_2O
c. SO_3
d. $SiBr_4$
e. $Co(OH)_2$
f. SF_6

10. **a.** carbon monoxide
b. boron trichloride
c. carbon disulfide
d. carbon tetrachloride
e. silicon dioxide
f. phosphorus triiodide
g. barium hydroxide
h. hydrogen borate

Answers to Chapter 2 Review Questions

1. c
2. c
3. a
4. d
5. b
6. e
7. b
8. d

17. **a.** 1:2
b. 1:1
c. 1:1
d. 1:1

19. **a.** magnesium chloride
b. sodium oxide
c. iron(III) chloride
d. copper(II) oxide
e. barium hypochlorite
f. ammonium nitrate
g. aqueous hydrogen chromate, or chromic acid
h. hydrogen phosphate
i. potassium hydroxide
j. cadmium hydroxide

20. **a.** $AuCl_3$
b. MgO
c. $LiNO_2$
d. Ca_3P_2
e. MnS
f. $Ca(ClO)_2$
g. HCl(aq)
h. $H_2SO_4(aq)$
i. $Co(OH)_2$
j. LiOH

22. **a.** sulfur dioxide
b. dinitrogen tetroxide
c. carbon monoxide
d. dichlorine oxide

23. **a.** H_2O
b. SO_3
c. $SiCl_4$

Answers to Chapter 2 Self Assessment Questions

1. b
2. e
3. b
4. a
5. e
6. d
7. b
8. c
9. b
10. c

15. **a.** $\Delta EN = 1.4$; polar covalent
b. $\Delta EN = 0.4$; slightly polar covalent
c. $\Delta EN = 0.0$; non-polar covalent
d. $\Delta EN = 1.8$; mostly ionic

16. **a.** magnesium phosphate
b. sodium iodate
c. aluminum phosphate
d. sodium hydrogen carbonate

17. **a.** KSCN
b. YCl_3
c. Fe_2S_3
d. SnF_2

18. **a.** trisilicon tetranitride
b. phosphorus pentachloride
c. sulfur hexafluoride
d. chlorine trifluoride

19. **a.** SO_3
b. CO
c. Se_2Br_2
d. NI

Answers to Unit 1 Review Questions

1. d
2. c
3. d
4. d
5. c
6. a
7. e
8. d
9. b
10. c

15. O, C, Ge, Ca, Ba

16. Cl, P, Mg, Ca, K

20 **a.** $\Delta EN = 2.1$; mostly ionic
b. $\Delta EN = 0.8$; polar covalent
c. $\Delta EN = 0$; non-polar covalent

28 **a.** Ca_3N_2
b. SI_2
c. $PbBr_2$
d. $AlPO_3$

29 **a.** nitrogen trichloride
b. potassium carbonate
c. iron(II) oxide
d. dinitrogen tetroxide

30 **a.** silicon
b. sulfur
c. phosphorus

31 **a.** nitric acid, aqueous hydrogen nitrate
b. aqueous hydrogen iodide, hydroiodic acid
c. aqueous hydrogen oxalate, oxalic acid
d. cobalt(III) hydroxide

32 **a.** HClO(aq)
b. NH_4OH

- **c.** HNO_2 (aq)
- **d.** $Mg(OH)_2$

34 a. ionic
- **b.** 2, 7, 7
- **c.** 0, 8, 8

35 a. 2+
- **b.** 2–; 0
- **c.** 1–; 0

57. 107.9 u

58. a. 3– charge
- **b.** 2– charge
- **c.** 2+ charge
- **d.** 1+ charge

59. a. one aluminum to three chlorine
- **b.** one aluminum to one nitrogen
- **c.** two aluminum to three oxygen

61. NH_3

Answers to Unit 1 Self-Assessment Questions

1. a
2. b
3. a
4. c
5. d
6. e
7. d
8. d
9. b
10. a
11. 3; 15
13. 107.9 u

18. a. lose 2 electrons
- **b.** gain 2 electrons
- **c.** lose 1 electron
- **d.** gain 3 electrons

19 a. ionic; Mg_3N_2
- **b.** covalent; OF_2
- **c.** ionic; $SnBr_2$
- **d.** ionic; $AlPO_4$
- **e.** ionic; $Co_2(SO_3)_3$

20 a. covalent; phosphorus pentachloride
- **b.** ionic; lithium carbonate
- **c.** ionic; copper(II) oxide
- **d.** covalent; dinitrogen trioxide
- **e.** ionic; ammonium nitrite

UNIT 2

Chapter 3

Answers to Learning Check Questions

1. A skeleton equation is a representation of a chemical reaction that does not include the relative quantities of the substances involved.

2. (aq)

3. a. The skeleton equation would show that carbon dioxide and water are the reactants and sugar and oxygen are the products; the equation would show the chemical formula and state for each substance.
- **b.** The skeleton equation would not show the relative number of atoms, molecules, or ions involved in the reaction.

4. The two-way arrow indicates that the reaction is reversible.

5. a. The state of water should be (ℓ) but is shown as (g). The second plus sign (between water and sodium hydroxide) should be an arrow. The formula for sodium hydroxide should be NaOH. The state of sodium hydroxide should be (aq). The formula for hydrogen should have a subscript 2.
- **b.** $Na(s) + H_2O(\ell) \rightarrow NaOH(aq) + H_2(g)$

6. a. reactant: liquid water; products: hydrogen gas and oxygen gas
  $H_2O(\ell) \rightarrow H_2(g) + O_2(g)$
- **b.** reactants: chlorine gas and aqueous zinc bromide; products: liquid bromine and zinc chloride
  $Cl_2(g) + ZnBr_2(aq) \rightarrow Br_2(\ell) + ZnCl_2(s)$

7. The general form of a synthesis reaction is $A + B \rightarrow AB$

8. A binary ionic compound is formed.

9. a. A single product is formed from two or more reactants.
- **b.** $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

10. a. $CaCl_2(s)$
- **b.** $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$

11. Your graphic organizer should show that the first reaction, $S(s) + O_2(g) \rightarrow SO_2(g)$, involves two elements; the second reaction, $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$, involves a compound and an element; and the third reaction, $SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(aq)$, involves two compounds.

12. No, there must be another product that contains hydrogen. Because there is more than one product, the reaction cannot be a synthesis reaction.

13. A metal hydroxide, or base, forms when a metal oxide reacts with water.

14. The graphic organizer should show that the solution formed from a metal oxide is basic and that the solution formed from a non-metal oxide is acidic.

15. $AB \rightarrow A + B$; a single reactant breaks down into two or more products.

16. No, the reactant must be a compound because chemical reactions can break a compound into simpler substances but cannot change elements into simpler substances.

17. Electrolysis is possible in the aqueous state. An aqueous solution of an ionic compound can conduct an electric current because when the compound dissolves in water, the ions separate and are free to move.

18. Thermal decomposition can be used to isolate elemental mercury by heating solid mercury(II) oxide, according to the following equation: $2HgO(s) \rightarrow 2Hg(\ell) + O_2(g)$ Uses of mercury include (any two of the following) thermometers, barometers, and dental fillings.

19. Combustion reactions release light and heat.

20. A hydrocarbon is a compound composed of only carbon and hydrogen.
- **a.** Hydrocarbons can take part in complete or incomplete combustion reactions.
- **b.** The products of complete combustion are carbon dioxide and water.

21. a. The reaction of nitrogen gas and oxygen gas is not a combustion reaction because energy in the form of heat is absorbed, not released.
- **b.** It is a synthesis reaction.

22. a. A blue flame indicates complete combustion is occurring, and a yellow flame indicates incomplete combustion.

b. The blue flame indicates complete combustion, so the flame is generating much more heat than light.

c. In **Figure 3.23**, carbon dioxide and water are formed. In **Figure 3.24**, carbon dioxide, water, elemental carbon, and possibly carbon monoxide are produced.

23. A gas stove is designed to be hot, so it should allow complete combustion to occur. In addition, because gas stoves are used indoors, it is important for the combustion to be complete to avoid the production of toxic carbon monoxide.

24. A combustion reaction requires oxygen, which is not present in space. The antenna array could not burn.

25. Incomplete combustion is a chemical reaction in which a substance reacts with oxygen but there is too little oxygen for complete combustion to occur. In addition to the carbon dioxide and water that are produced during complete combustion reactions, elemental carbon and carbon monoxide are produced during incomplete combustion reactions.

26. Carbon monoxide, a poisonous gas, is formed during incomplete combustion.

27. Carbon monoxide production can occur in pulp and paper production, petroleum refineries, and steel production.

28. The amount of oxygen present determines whether complete or incomplete combustion will occur.

29. If the synthesis reaction involves an element or compound reacting with oxygen, it can also be classified as a combustion reaction.

30. This process is cellular respiration, because it occurs at fairly low temperatures and is mediated by living organisms.

Answers to Practice Problems

1. $H_2(g) + O_2(g) \rightarrow H_2O(g)$

2. $Na(s) + H_2O(\ell) \rightarrow NaOH(aq) + H_2(g)$

3. $KClO_3(s) \rightarrow KCl(s) + O_2(g)$

4. $Cu(s) + O_2(g) \rightarrow CuO(s)$

5. $AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

6. $C_3H_8(g) + O_2(g) \rightarrow H_2O(g) + CO_2(g)$

7. $SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(aq)$

8. $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$

9. $AlF_3(s) \rightarrow Al(s) + F_2(g)$

10. $Hg(\ell) + O_2(g) \rightarrow HgO(s)$

11. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

12. $3Mg(s) + 2AlCl_3(aq) \rightarrow 2Al(s) + 3MgCl_2(aq)$

13. $2NaOH(aq) + CuCl_2\ (aq) \rightarrow 2NaCl(aq) + Cu(OH)_2(s)$

14. $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

15. $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

16. $4Al(s) + 3MnO_2(s) \rightarrow 2Al_2O_3(s) + 3Mn(s)$

17. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

18. $4NH_3(g) + 7O_2(g) \rightarrow 4NO_2(g) + 6H_2O(\ell)$

19. $K_2S(aq) + CoCl_2(aq) \rightarrow 2KCl(aq) + CoS(s)$

20. $2HCl(g) + Na_2CO_3(aq) \rightarrow CO_2(g) + H_2O(\ell) + 2NaCl(aq)$

21. lithium oxide; $4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$

22. strontium fluoride; $Sr(s) + F_2(g) \rightarrow SrF_2(s)$

23. iron(II) bromide or iron(III) bromide; $2Fe(s) + 3Br_2(\ell) \rightarrow 2FeBr_3(s)$ or $Fe(s) + Br_2(\ell) \rightarrow FeBr_2(s)$

24. phosphorus trihydride; $2P(s) + 3H_2(g) \rightarrow 2PH_3(g)$

25. calcium iodide; $Ca(s) + I_2(s) \rightarrow CaI_2(s)$

26. tin(IV) oxide or tin(II) oxide; $Sn(s) + O_2(g) \rightarrow SnO_2(s)$ or $2Sn(s) + O_2(g) \rightarrow 2SnO(s)$

27. bismuth(III) sulfide or bismuth(V) sulfide; $2Bi(s) + 3S(s) \rightarrow Bi_2S_3(s)$ or $2Bi(s) + 5S(s) \rightarrow Bi_2S_5(s)$

28. aluminum iodide; $2Al(s) + 3I_2(s) \rightarrow 2AlI_3(s)$

29. silver oxide; $4Ag(s) + O_2(g) \rightarrow 2Ag_2O(s)$

30. nitrogen dioxide; $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$

31. potassium and bromine; $2KBr(\ell) \rightarrow 2K(\ell) + Br_2(\ell)$

32. aluminum and oxygen; $2Al_2O_3(\ell) \rightarrow 4Al(\ell) + 3O_2(g)$

33. magnesium oxide and water; $Mg(OH)_2(s) \rightarrow MgO(s) + H_2O(g)$

34. calcium nitrite and oxygen; $Ca(NO_3)_2(s) \rightarrow Ca(NO_2)_2(s) + O_2(g)$

35. copper(II) oxide and carbon dioxide; $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

36. chromium and chlorine; $2CrCl_3(\ell) \rightarrow 2Cr(\ell) + 3Cl_2(g)$

37. barium oxide and carbon dioxide; $BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$

38. rubidium nitrite and oxygen; $2RbNO_3(s) \rightarrow 2RbNO_2(s) + O_2(g)$

39. lithium oxide and water; $2LiOH(s) \rightarrow Li_2O(s) + H_2O(g)$

40. magnesium and chlorine; $MgCl_2(s) \rightarrow Mg(s) + Cl_2(g)$

41. $C_7H_{16}(\ell) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(g)$

42. $C_9H_{20}(\ell) + 14O_2(g) \rightarrow 9CO_2(g) + 10H_2O(g)$

43. $2C_2H_2(\ell) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$

44. $2C_6H_6(\ell) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$

45. $2C_8H_{18}(\ell) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

46. $2C_8H_{18}(\ell) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(g)$

47. $2C_5H_{12}(\ell) + 11O_2(g) \rightarrow 10CO(g) + 12H_2O(g)$

48. $C_3H_8(g) + 2O_2(g) \rightarrow 3C(s) + 4H_2O(g)$

49. $4C_7H_{16}(\ell) + 37O_2(g) \rightarrow 14CO_2(g) + 14CO(g) + 32H_2O(g)$

50. a. $C_6H_{12}(\ell) + 6O_2(g) \rightarrow 6CO(g) + 6H_2O(g)$

b. $C_6H_{12}(\ell) + 3O_2(g) \rightarrow 6C(s) + 6H_2O(g)$

Answers to Caption Questions

Figure 3.4: There is one aluminum atom in both the reactants and the product; however, there are two bromine atoms in the reactants and three in the product.

Figure 3.5: In the reactants, there are three molecules of bromine, each containing two atoms for a total of six bromine atoms. In the products, there are two formula units of aluminum bromide, each containing three bromine atoms, for a total of six. The number of bromine atoms in the reactants and products is balanced, with six on each side.

Figure 3.8: Both the general form of a synthesis reaction and the reaction shown illustrate two separate substances joining together to form one substance.

Figure 3.10: Manganese and copper are both multivalent metals. Like copper, manganese can form different binary compounds depending on the reaction that occurs.

Figure 3.14: The nitrogen gas causes the air bag to inflate.

Figure 3.24: Elemental carbon forms dark soot deposits on surfaces.

Figure 3.25: This reaction is a combustion reaction because hydrogen reacts with oxygen to form an oxide, and noticeable heat and light are produced.

Figure 3.26: Each product is an oxide of the element.

Answers to Section 3.1 Review Questions

7. (g)

11. a. $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
  b. $2Al(s) + 3CuSO_4(aq) \rightarrow 3Cu(s) + Al_2(SO_4)_3(aq)$
  c. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
  d. $CaCl_2(aq) + F_2(g) \rightarrow CaF_2(aq) + Cl_2(g)$

12. a. hydroxide, OH^-, and phosphate, PO_4^{3-}
  c. $Ba(OH)_2(s) + H_3PO_4(aq) \rightarrow BaHPO_4(s) + 2H_2O(\ell)$

13. The coefficients are not in the lowest possible ratio. The correct equation is: $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(\ell)$.

14. The formulas are incorrectly written. Therefore, the equation is not balanced properly. $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

Answers to Section 3.2 Review Questions

5. $2Al(s) + 3S(s) \rightarrow Al_2S_3(s)$

15. lithium carbonate → lithium oxide + carbon dioxide; lithium hydroxide → lithium oxide + water

Answers to Section 3.3 Review Questions

5. $5CO_2(g)$

6. a. $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
  b. $2C_{10}H_{22}(\ell) + 31O_2(g) \rightarrow 20CO_2(g)) + 22H_2O(g)$
  c. $C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$
  d. $2C_6H_{14}(\ell) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

Answers to Chapter 3 Review Questions

1. c **3.** e **5.** c **7.** d
2. a **4.** c **6.** a **8.** e

10. 1

18. a. $Mg_3N_2(s) \rightarrow 3Mg(s) + N_2(g)$
  b. $4Mn(s) + 3O_2(g) \rightarrow 2Mn_2O_3(s)$
  c. $CO_2(g) + 4H_2(g) \rightarrow CH_4(g) + 2H_2O(g)$
  d. $2PbO(s) \rightarrow 2Pb(s) + O_2(g)$
  e. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$
  f. $Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$
  g. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
  h. $3PbC_{l4}(aq) + 4K_3PO_4(aq) \rightarrow 12KCl(aq) + Pb_3(PO_4)_4(s)$

20. a. potassium sulfide; $2K(s) + S(s) \rightarrow K_2S(s)$
  b. chromium(III) chloride; $2Cr(s) + 3Cl_2(g) \rightarrow 2CrCl_3(s)$; chromium(II) chloride; $Cr(s) + Cl_2(g) \rightarrow CrCl_2(s)$
  c. silver oxide; $4Ag(s) + O_2(g) \rightarrow 2Ag_2O(s)$
  d. sulfur hexachloride; $S(s) + 3Cl_2(g) \rightarrow SCl_6(g)$

21. a. magnesium and iodine; $MgI_2(\ell) \rightarrow Mg(s) + I_2(g)$
  b. copper(II) nitrite and oxygen; $Cu(NO_3)_2(s) \rightarrow Cu(NO_2)_2(s) + O_2(g)$
  c. barium oxide and carbon dioxide; $BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$

23. a. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$
  b. $C_5H_{12}(\ell) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$
  c. $2C_8H_{18}(\ell) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

27. Your diagram should include the following chemical equations: $Ca(s) + Br_2(\ell) \rightarrow CaBr_2(s)$; $Mg(s) + Br_2(\ell) \rightarrow MgBr_2(s)$; $Sr(s) + Br_2(\ell) \rightarrow SrBr_2(s)$

31. The chemical formulas for nitrogen gas and hydrogen gas are incorrect, so the chemical equation is not correctly balanced; in addition, the arrow should show that the chemical reaction is reversible. The correct chemical equation is $N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$.

34. b. $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$

35. a. $2H_2O_2(\ell) \rightarrow 2H_2O(\ell) + O_2(g)$

37. a. decomposition; $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
  b. synthesis; $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$
  c. synthesis; $2CaSO_3(s) + O_2(g) \rightarrow 2CaSO_4(s)$

Answers to Chapter 3 Self-Assessment Questions

1. e **3.** c **5.** d **7.** a **9.** b
2. b **4.** e **6.** a **8.** d **10.** b

11. $CO_2(g)$

12. a. $Na_2CO_3(s) \rightarrow Na_2O(s) + CO_2(g)$

15. a. $Cr(ClO_3)_2(s) \rightarrow CrCl_2(s) + 3O_2(g)$
  b. $4Rb(s) + O_2(g) \rightarrow 2Rb_2O(s)$
  c. $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
  d. $2KOH(s) \rightarrow K_2O(s) + H_2O(g)$

16. a. combustion; $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
  b. decomposition; $2KBrO_3(s) \rightarrow 2KBr(s) + 3O_2(g)$
  c. synthesis; $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$
  d. decomposition; $Ca(NO_3)_2(s) \rightarrow Ca(NO_2)_2(s) + O_2(g)$
  e. decomposition; $C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(\ell)$
  f. combustion; $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

17. a. aluminum chloride; $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$
  b. barium hydroxide; $BaO(s) + H_2O(\ell) \rightarrow Ba(OH)_2(s)$

18. a. calcium and nitrogen; $Ca_3N_2(s) \rightarrow 3Ca(s) + N_2(g)$
  b. sulfur dioxide and water; $H_2SO_3(aq) \rightarrow SO_2(g) + H_2O(\ell)$

19. a. $C_5H_{12}(\ell) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$
  b. $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

Chapter 4

Answers to Learning Check Questions

1. $A + BX \rightarrow AX + B$

2. An element displaces a different element from a compound, forming a new compound and the replaced element as products.

3. Scientists perform experiments to determine the relative reactivity of a series of elements.

4. Platinum and gold, which are at the bottom of the reactivity series of metals, are the least reactive metals. A platinum or gold coating can prevent another, more reactive, metal underneath it from reacting with any substances the metal object might come into contact with.

5. **a.** No reaction.
   b. A reaction would not occur because copper is not reactive enough to displace lead.

6. **a.** Titanium would appear closer to the bottom of the series.
   b. Its use in medical implants indicates that titanium is not very reactive.

7. AX + BY → AY + BX

8. They are in aqueous solution.

9. Cations are positive ions, so describing a double displacement reaction as the exchange of cations is correct.

10. The ions in each reactant switch partners, so knowing which ions are involved makes it possible to correctly pair them up and determine the reaction products.

11. No, the products of a double displacement reaction are generally two compounds, not elements.

12. **a.** potassium nitrate and silver bromide; $KNO_3(aq)$ and $AgBr(s)$
    b. $KBr(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgBr(s)$

13. A precipitate is an insoluble solid that forms as a result of a chemical reaction between two soluble compounds.

14. The other compounds are in aqueous solution, but the silver chloride is a solid precipitate.

15. No, the solubility rules in the table are for the solubility of compounds in water only.

16. A double displacement reaction that produces a gas can produce either carbon dioxide or ammonia. The general forms of these reactions are as follows: acid + compound containing carbonate ion → ionic compound + water + carbon dioxide; compound containing ammonium ions + compound containing hydroxide ions → ionic compound + water + ammonia

17. **a.** $Ba(OH)_2(s)$
    b. $MgS(s)$
    c. $H_3PO_4(s)$
    d. $Na_2SO_3(s)$

18. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$

19. Single displacement; aluminum displaces iron from iron oxide.

20. The liquid metal product of a thermite reaction is useful for welding.

21. Solid aluminum reacts with solid copper(II) oxide to produce pure liquid copper and aluminum oxide.

22. Seashells are a source of calcium carbonate, which is decomposed to produce calcium oxide. Calcium oxide is the reactant in the next step in the process of magnesium extraction.

23. The precipitation allows for the magnesium to be easily separated from the other ions in seawater.

24. Chlorine is re-used to produce hydrochloric acid, which is needed for the neutralization reaction.

Answers to Practice Problem Questions

1. $Mg(s) + CrSO_4(aq) \rightarrow MgSO_4(aq) + Cr(s)$
2. NR
3. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
4. $F_2(g) + MgI_2(aq) \rightarrow MgF_2(aq) + I_2(s)$
5. $Cl_2(g) + 2NaI(aq) \rightarrow 2NaCl(aq) + I_2(s)$
6. NR
7. NR
8. $2K(s) + 2H_2O(\ell) \rightarrow 2KOH(aq) + H_2(g)$
9. $2HCl(aq) + Cd(s) \rightarrow H_2(g) + CdCl_2(aq)$
10. $3Pb(ClO_3)_4(aq) + 4Al(s) \rightarrow 4Al(ClO_3)_3(aq) + 3Pb(s)$
11. potassium chloride and calcium sulfate; $K_2SO_4(aq) + CaCl_2(aq) \rightarrow 2KCl(aq) + CaSO_4(s)$
12. barium carbonate and sodium nitrate; $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaNO_3(aq)$
13. iron(III) hydroxide and sodium chloride; $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$
14. rubidium iodide and copper(II) sulfide; $Rb_2S(aq) + CuI_2(aq) \rightarrow 2RbI(aq) + CuS(s)$
15. zinc acetate and copper(I) bromide; $ZnBr_2(aq) + 2CuCH_3COO(aq) \rightarrow Zn(CH_3COO)_2(aq) + 2CuBr(s)$
16. lithium chloride and magnesium hydroxide; $2LiOH(aq) + MgCl_2(aq) \rightarrow 2LiCl(aq) + Mg(OH)_2(s)$
17. aluminum nitrate and lead(II) sulfate; $Al_2(SO_4)_3(aq) + 3Pb(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3PbSO_4(s)$
18. lithium chloride and magnesium phosphate; $2Li_3PO_4(aq) + 3MgCl_2(aq) \rightarrow 6LiCl(aq) + Mg_3(PO_4)_2(s)$
19. calcium sulfate and magnesium nitrate; $Ca(NO_3)_2(aq) + MgSO_4(aq) \rightarrow CaSO_4(s) + Mg(NO_3)_2(aq)$
20. silver chloride and magnesium nitrate; $2AgNO_3(aq) + MgCl_2(aq) \rightarrow 2AgCl(s) + Mg(NO_3)_2(aq)$
21. potassium chloride, water, and carbon dioxide; $K_2CO_3(aq) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2O(\ell) + CO_2(g)$
22. sodium sulfate, water, and carbon dioxide; $H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow Na_2SO_4(aq) + H_2O(\ell) + CO_2(g)$
23. sodium chloride, water, and ammonia; $NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(\ell) + NH_3(g)$
24. rubidium chloride and water; $RbOH(aq) + HCl(aq) \rightarrow RbCl(aq) + H_2O(\ell)$
25. calcium acetate, water, and carbon dioxide; $CaCO_3(s) + 2HCH_3COO(aq) \rightarrow Ca(CH_3COO)_2(aq) + H_2O(\ell) + CO_2(g)$
26. lithium bromide, water, and ammonia; $LiOH(aq) + NH_4Br(aq) \rightarrow LiBr(aq) + H_2O(\ell) + NH_3(g)$
27. lithium sulfate and water; $H_2SO_4(aq) + 2LiOH(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(\ell)$
28. lithium acetate, water, and carbon dioxide; $LiHCO_3(aq) + HCH_3COO(aq) \rightarrow LiCH_3COO(aq) + H_2O(\ell) + CO_2(g)$
29. calcium nitrate and water; $Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(\ell)$
30. magnesium chloride, water, and ammonia; $2NH_4Cl(aq) + Mg(OH)_2(aq) \rightarrow MgCl_2(aq) + 2H_2O(\ell) + 2NH_3(g)$

Answers to Caption Questions

Figure 4.2: Nothing. The nitrate ions do not change during the reaction.

Figure 4.5: Metals that can displace hydrogen from acids are tin, nickel, cobalt, cadmium, iron, chromium, zinc, aluminum, sodium, calcium, barium, potassium, and lithium; metals that cannot displace hydrogen from acids are copper, silver, mercury, platinum, and gold.

Figure 4.9: The positive ions, Ag^+ and Na^+, change places.

Figure 4.13: The thermite reaction occurs in the solid state, but most double replacement reactions occur in an aqueous solution.

Figure 4.16: Aluminum-magnesium tubing is strong, light, and more resistant to corrosion than pure aluminum—all of which are valuable properties for a kayak.

Figure 4.17: The furnaces are built at different elevations, so the material can flow downhill, moving by gravity from one furnace to the next.

Figure 4.19: A precipitate holds the cyanide in place as a solid, increasing the chance that it can be cleaned up before it is washed into groundwater or a river.

Answers to Section 4.1 Review Questions

5. a. NR
  b. A reaction occurs. Zinc displaces iron. $Zn(s) + FeCl_2(aq) \rightarrow Fe(s) + ZnCl_2(aq)$
  c. A reaction occurs. Magnesium displaces aluminum. $3Mg(s) + Al_2(SO_4)_3(aq) \rightarrow 2Al(s) + 3MgSO_4(aq)$
  d. A reaction occurs. Zinc displaces hydrogen. $Zn(s) + 2HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$
  e. NR
  f. A reaction occurs. Magnesium displaces hydrogen. $Mg(s) + H_2SO_4(aq) \rightarrow H_2(g) + MgSO_4(aq)$

12. If the liquid were water, the metal shown could not be zinc, because zinc does not react with water to form hydrogen gas. If the liquid were an acid, then the metal could be zinc, because zinc can displace hydrogen from an acid.

15. a. A reaction occurs. Iron displaces hydrogen. $Fe(s) + 2HBr(aq) \rightarrow H_2(g) + FeBr_2(s)$ or $2Fe(s) + 6HBr(aq) \rightarrow 3H_2(g) + 2FeBr_3(s)$
  b. A reaction occurs. Bromine displaces iodine. $Br_2(\ell) + MgI_2(aq) \rightarrow MgBr_2(aq) + I_2(aq)$
  c. A reaction occurs. Magnesium displaces aluminum. $3Mg(s) + Al_2(SO_4)_3(aq) \rightarrow 2Al(s) + 3Mg(SO_4)(aq)$
  d. A reaction occurs. Lithium displaces hydrogen. $2Li(s) + 2H_2O(\ell) \rightarrow 2LiOH(aq) + H_2(g)$
  e. NR
  f. NR

Answers to Section 4.2 Review Questions

1. b. CW + DZ → CZ + DW

9. b. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$

11. d. $2NaOH(aq) + CuCl_2(aq) \rightarrow 2NaCl(aq) + Cu(OH)_2(s)$

12. ammonia, $NH_3(g)$; $NH_4Br(aq) + NaOH(aq) \rightarrow NaBr(aq) + H_2O(\ell) + NH_3(g)$

15. calcium chloride and water; $2HCl(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(\ell) + CaCl_2(aq)$

16. Although the chemical formulas are correct, the equation is not correctly balanced because water needs a coefficient. The states of the products are incorrect. Water should be liquid, and sodium phosphate should be in aqueous solution. Correct equation: $3NaOH(aq) + H_3PO_4(aq) \rightarrow 3H_2O(\ell) + Na_3PO_4(aq)$

Answers to Section 4.3 Review Questions

8. b. $MgCl_2(\ell) \rightarrow Mg(\ell) + Cl_2(g)$

13. c. $2CuFeS_2(s) + 4O_2(g) \rightarrow Cu_2S(\ell) + 2FeO(\ell) + 3SO_2(g)$
$2Cu_2S(\ell) + 3O_2(g) \rightarrow 2Cu_2O(\ell) + 2SO_2(g)$
$Cu_2S(\ell) + 2Cu_2O(\ell) \rightarrow 6Cu(\ell) + SO_2(g)$

Answers to Chapter 4 Review Questions

1. c
2. a
3. b
4. d
5. c
6. d
7. d
8. b

15. A + BX → BA + X

20. a. $3Mg(s) + 2Co(NO_3)_3(aq) \rightarrow 3Mg(NO_3)_2(aq) + 2Co(s)$
  b. $Cl_2(g) + 2LiBr(aq) \rightarrow Br_2(g) + 2LiCl(aq)$
  c. $Zn(s) + 2HClO_4(aq) \rightarrow Zn(ClO_4)_2(aq) + H_2(g)$
  d. NR
  e. $2Al(s) + 3NiCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Ni(s)$
  f. $2K(s) + 2H_2O(\ell) \rightarrow 2KOH(aq) + H_2(g)$
  g. NR

23. a. potassium bromide and barium sulfate; $K_2SO_4(aq) + BaBr_2(aq) \rightarrow 2KBr(aq) + BaSO_4(s)$
  b. lithium nitrate, water, and carbon dioxide; $2HNO_3(aq) + Li_2CO_3(aq) \rightarrow 2LiNO_3(aq) + H_2O(\ell) + CO_2(g)$
  c. copper(II) hydroxide and sodium bromide; $CuBr_2(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + 2NaBr(aq)$
  d. rubidium nitrate and lead(II) sulfide; $Rb_2S(aq) + Pb(NO_3)_2(aq) \rightarrow 2RbNO_3(aq) + PbS(s)$
  e. potassium sulfate, water, and ammonia; $(NH_4)_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(\ell) + 2NH_3(g)$
  f. iron(II) nitrate and silver bromide; $FeBr_2(aq) + 2AgNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + 2AgBr(s)$
  g. lithium sulfate and water; $2LiOH(aq) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(\ell)$

25. $2Al(s) + 3H_2SO_4(aq) \rightarrow 3H_2(g) + Al_2(SO_4)_3(aq)$;
$2Al(OH)_3(aq) + 3H_2SO_4(aq) \rightarrow 6H_2O(\ell) + Al_2(SO_4)_3(aq)$

26. Ammonia gas can form from a double displacement reaction if an ammonium compound and a hydroxide react together. The ammonia is produced when ammonium hydroxide formed in the double displacement decomposes. $NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + NH_4OH(aq)$ followed by $NH_4OH(aq) \rightarrow H_2O(\ell) + NH_3(g)$

38. b. React magnesium metal with hydrochloric acid: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$. Add sodium phosphate solution to the magnesium chloride solution formed: $3MgCl_2(aq) + 2Na_3PO_4(aq) \rightarrow Mg_3(PO_4)_2(s) + 6NaCl(aq)$. Filter the resulting products to collect the magnesium phosphate precipitate.

40. $HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2O(\ell) + CO_2(g)$

Answers to Chapter 4 Self-Assessment Questions

1. e
2. b
3. e
4. d
5. b
6. d
7. b
8. a
9. c
10. d

12. a. silver chloride and lithium nitrate; $LiCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + LiNO_3(aq)$

b. NR

c. potassium chloride and iodine; $Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(s)$

d. lead(II) sulfate and sodium nitrate; $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$

14. a. A precipitate will form if calcium chloride or lead(II) acetate is present.

b. $CaCl_2(aq) + Na_2SO_4(aq) \rightarrow CaSO_4(s) + 2NaCl(aq)$; $Pb(CH_3COO)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaCH_3COO(aq)$

15. a. Water should be a liquid. The product should be potassium hydroxide rather than potassium oxide. $2K(s) + 2H_2O(\ell) \rightarrow 2KOH(aq) + H_2(g)$

b. Lithium nitrate is soluble, so it should be marked aqueous rather than solid, and lead(II) chloride is not soluble, so it should be marked solid rather than aqueous. A coefficient is needed in front of lithium nitrate to balance the equation. $2LiCl(aq) + Pb(NO_3)_2(aq) \rightarrow 2LiNO_3(aq) + PbCl_2(s)$

23. a. $H_2CO_3(aq) + 2NaOH(aq) \rightarrow Na_2CO_3(aq) + 2H_2O(\ell)$

c. $Na_2CO_3(aq) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + 2NaOH(aq)$

Answers to Unit 2 Review Questions

1. d

2. d

3. c

4. c

5. b

6. e

7. b

8. c

9. e

10. b

11. $3NaOH(aq) + AlCl_3(aq) \rightarrow 3NaCl(aq) + Al(OH)_3(aq)$

16. a. $Ca_3N_2(s) \rightarrow 3Ca(s) + N_2(g)$

b. $4Cr(s) + 3O_2(g) \rightarrow 2Cr_2O_3(s)$

c. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

d. $2BaO(s) \rightarrow 2Ba(s) + O_2(g)$

26. b. The solution is acidic and will turn blue litmus paper red. $CO_2(g) + H_2O(\ell) \rightarrow H_2CO_3(aq)$

27. c

28. a. $2C_6H_6(\ell) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(\ell)$

b. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

c. $Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(s) + Br_2(g)$

d. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

30. a. $Cu(NO_3)_2(aq)$; copper(II) nitrate

b. $Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$

c. 188 g

31. d. $HCl(aq) + AgNO_3(aq) \rightarrow HNO_3(aq) + AgCl(s)$

32. a. NR. Neither of the possible products will form a precipitate.

b. $Nb_2(SO_4)_5(aq) + 5Ba(NO_3)_2(aq) \rightarrow 2Nb(NO_3)_5(aq) + 5BaSO_4(s)$

c. $SrBr_2(aq) + 2AgNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + 2AgBr(s)$

45. Sample answer: The reactivity of sodium is too great for this process. You are counting on the reaction $Na(s) + AgNO_3(aq) \rightarrow NaNO_3(aq) + Ag(s)$. However, sodium is reactive enough to displace hydrogen from water by the reaction $2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g)$. Besides the dangerous nature of the sodium itself, the sodium hydroxide is caustic, and the hydrogen is flammable.

48. b. $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$; $Mg(OH)_2(s) \rightarrow MgO(s) + H_2O(g)$

50. a. $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(\ell) + CO_2(g)$

52. a. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

53. formula for DTBP: $C_8H_{18}O_2(g)$ or $(CH_3)_3COOC(CH_3)_3(g)$; complete combustion: $2C_8H_{18}O_2(g) + 23O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$ A fuel that could power an engine in the absence of oxygen could be used in low-oxygen environments, such as to power a chainsaw for use by a firefighter within a burning building.

54. a. $Ni(s) + 4CO(g) \rightarrow Ni(CO)_4(g)$

55. b. $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(\ell + CO_2(g)$; $CuSO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + Cu(s)$

Answers to Unit 2 Self-Assessment Questions

1. d

2. b

3. e

4. b

6. c

7. b

8. c

9. a

10. a

12. a. $Br_2(\ell) + 2NaI(aq) \rightarrow 2NaBr(aq) + I_2(s)$

b. $2Al(s) + 3Cu(NO_3)_2(aq) \rightarrow 3Cu(s) + 2Al(NO_3)_3(aq)$

c. $2Fe_2O_3(s) \rightarrow 4Fe(s) + 3O_2(g)$

d. $Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(\ell)$

e. $6Li(s) + N_2(g) \rightarrow 2Li_3N(s)$

f. $2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$

15. a. $C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$

b. $2Al(s) + 3Br_2(\ell) \rightarrow 2AlBr_3(s)$

c. $2RbNO_3(s) \rightarrow 2RbNO_2(s) + O_2(g)$

16. a. $2LiCl(\ell) \rightarrow 2Li(\ell) + Cl_2(g)$

b. If water enters the reaction chamber, the lithium that forms might react with the water and produce flammable hydrogen gas, which could ignite and cause injury to the scientist. $2Li(\ell) + 2H_2O(\ell) \rightarrow 2LiOH(aq) + H_2(g)$

18. a. $SiO_2(s) + C(s) \rightarrow CO_2(g) + Si(\ell)$

20. b. $H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2O(\ell) + CO_2(g) + CaSO_4(aq)$

23. a. $MgS(aq) + Cu(NO_3)_2(aq) \rightarrow CuS(s) + Mg(NO_3)_2(aq)$

b. No reaction

c. $Br_2(g) + 2KI(aq) \rightarrow 2KBr(aq) + I_2(s)$

d. No reaction

Unit 3

Chapter 5

Answers to Caption Questions

Figure 5.1: A score is 20, a gross is 12 × 12 or 144, a great gross is twelve gross or 1728, a baker's dozen is 13, a paper bale is 5000 sheets of paper.

Figure 5.5: 6.02×10^{22} molecules

Figure 5.10: 0.495 mol

Answers to Learning Check Questions

1. Avogadro's constant is defined as exactly equal to the number of atoms of carbon-12 in 1 g of carbon-12. However, the numerical value must be determined experimentally. Scientists are constantly updating the value as they improve the methods used to determine the value.

2. A mole is the mass of particles (atoms, ions, molecules, or formula units) that contain the same number of particles as 12 g of the isotope carbon-12. The numerical value of the number of particles that constitute a mole is the same as the Avogadro constant.

3. You would have two times the Avogadro constant of hydrogen atoms. Rounded off, the number would be $2(6.02 \times 10^{23}) = 1.20 \times 10^{24}$ hydrogen atoms.

4. You would not be able to see one person, but a mole of people is so many that they would be visible, as a group, from space. In fact, a mole of people would have a mass about the same size as the mass of Earth.

5. four

6. Paper is often purchased in large quantities because people use paper in large quantities. If it was purchased by the dozen, people would have to calculate how many dozen sheets of paper would be enough to use, so measuring paper by the dozen is an inefficient measurement.

7. Oxygen was originally chosen by chemists as the reference for measuring atomic masses. However, with the development of the mass spectrometer, the reference was changed to carbon. Mass spectrometers accelerate particles in a vacuum, through a magnetic field, which causes a deflection in the path of the particles. A reference mass is needed but oxygen not practical. To create a vacuum, strong pumps must be used. A tiny amount of the carbon in the lubricants for the pumps always got into the vacuum. Most naturally occurring carbon is carbon-12. So, since it was already there, chemists decided to use carbon-12 as the universal reference for atomic masses.

8. A mole is an exceptionally huge number. It was designed specifically to apply to atoms and molecules. To try to use a mole for measuring objects larger than atoms and molecules is exceedingly impractical. All values would be expressed in minute fractions of a mole. For example, a dozen would be about 2×10^{-23} mol. A great gross would be 2.87×10^{-21} mol. It defies common sense.

9. Atomic molar mass refers to the mass of one mole of atoms, whereas molar mass is more general referring to the mass of one mole of any entity, such as atoms, molecules, formula units, etc.

10. The numerical value for the atomic mass in both units is the same.

11. 63.55 g

12. This value can also be described as the atomic molar mass, or just the molar mass. The units are g/mol, meaning the mass of one mole of atoms of a certain element.

13. The reported values of atomic mass are the weighted averages of the naturally occurring isotopes. In addition, the actual value of the masses of individual isotopes measured by a mass spectrometer are not whole numbers.

Answers to Practice Problem Questions

1. 9×10^{13} refrigerators

2. 1.6×10^{27} km

3. 1.91×10^{16} y

4. 5.8×10^{16} km high

5. 2.1×10^{11} Rogers Centres

6. 2.48×10^{15} rows

7. One mole of tablespoons has a volume of 9.03×10^{9} km^3. Because this volume is greater than the total volume of the oceans, you would drain the oceans. In fact, you could drain the equivalent of over six times the world's oceans.

8. 1.91×10^{12} dollars/s

9. The Earth is 4.1×10^{3} times heavier

10. 8.3×10^{-17} cm

11. 6.38×10^{22} atoms

12. 5×10^{21} atoms

13. 5.1×10^{27} molecules

14. 5.11×10^{26} formula units

15. 3.15×10^{22} formula units

16. 2.32×10^{24} molecules

17. 1.3×10^{27} atoms

18. **a.** 2.90×10^{24} molecules
   b. 1.45×10^{25} atoms

19. **a.** 4.36×10^{25} atoms
   b. 2.18×10^{26} molecules

20. **a.** 0.015 mol
   b. 1.05×10^{23} atoms of C

21. 0.158 mol

22. 0.277 mol

23. 2.0×10^{2} mol

24. 1.4×10^{-4} mol

25. 5.1×10^{4} mol

26. 27.5 mol

27. 2.0 mol

28. 0.0346 mol

29. 1.3×10^{3} mol

30. 0.106 mol

31. **a.** 22.99 g/mol
   b. 183.84 g/mol
   c. 131.29 g/mol
   d. 58.69 g/mol

32. 123.88 g/mol

33. 310.18 g/mol

34. 391.2 g/mol

35. 113.94 g/mol

36. 306.52 g/mol

37. 315.51 g/mol

38. 133.98 g/mol

39. 392.21 g/mol

40. 132.91 g/mol

41. 182 g

42. 11 g

43. 0.231 g or 231 mg

44. 5.3×10^{2} mg

45. **a.** cobalt(II) nitrate 8.2×10^{-1} g
b. lead(IV) thiosulfate 1.28×10^4 g

46. **a.** NH_4NO_3 3.9×10^2 g

b. Fe_2O_3 2.59×10^3 g

47. 2.4×10^2 mg

48. 1.001 kg

49. **a.** Br_2
b. $Sr(IO_3)_2$

50. aluminum iodate

51. 1.73 mol

52. 139 mol

53. 8.75×10^{-4} mol

54. 1.1×10^{-4} mol

55. **a.** SiO_2, 6.2×10^{-5} mol
b. $Ti(NO_3)_4$, 0.08577 mol
c. $In_2(CO_3)_3$, 4.70×10^{-5} mol
d. 313 mol $CuSO_4 \cdot 5H_2O$,

56. 1.47 mol

57. 1.80×10^2 mol

58. 1.52×10^{-5} mol

59. $Al(OH)_3(s)$, $AgCl(s)$, $Ni(NO_3)_2(s)$

60. tin(IV) oxide, glucose, barium perchlorate

61. **a.** 3.52×10^3 g
b. 616 g
c. 7.00 g
d. 1.8×10^6 g

62. **a.** 2.40×10^{22} formula units
b. 4.80×10^{19} molecules
c. 9.36×10^{25} formula units
d. 1.75×10^{24} formula units

63. **a.** 9.52 g
b. 2.51×10^{24} formula units

64. **a.** 2.1×10^{18} molecules
b. 1.1 mg/day

65. 3.24×10^{-22} g

66. **a.** 2.58×10^{24} atoms
b. 6.88×10^{24} atoms
c. 8.60×10^{23} atoms

67. **a.** 1.26×10^{22} formula units
b. 6.29×10^{22} ions

68. beryllium arsenide

69. 0.37 g

70. $HCN(\ell)$, $CH_3COOH(\ell)$, $C_{12}H_{22}O_{11}(s)$

Answers to Section 5.1 Review Questions

4. 2.992×10^{-26} L

6. 3.48×10^{22} formula units of sodium chloride

7. **a.** 9.39×10^{22} atoms of gold
b. 4.7×10^{24} formula units of magnesium chloride
c. 9.15×10^{24} molecules of hydrogen peroxide

8. **a.** 4.5×10^{-4} mol
b. 0.0117 mol

9. the sample of carbon

10. 2.37×10^{-3} mol of water

11. **a.** 1.6×10^{24} formula units
b. 7.8×10^{24} atoms
c. 3.2×10^{24} atoms

12. **a.** 2.8×10^{23} formula units
b. 5.5×10^{23} atoms of chlorine

14. octane, then sodium hydrogen carbonate, then copper

Answers to Section 5.2 Review Questions

4. 0.241 mol

5. 1.0×10^5 g

6. 1.80×10^2 mol

7. 2.7×10^2 g

8. 5.2×10^4 g

9. 2.42×10^{23} atoms

10. aluminum sulfate

11. carbon dioxide

12. **a.** 2.1×10^{17}
b. 6.3×10^{17} atoms

13. 3.1×10^{17} lead atoms translates into 0.11 mg/L, which is about ten times greater than the allowable lead limit. Therefore, the water is not safe to drink.

14.

Substance	Number of Particles	Amount (mol)	Mass (g)
$P_4(s)$	7.95×10^{24}	13.2	1.64×10^3
$Ba(MnO_4)_2(s)$	6.7×10^{20}	1.1×10^{-3}	0.42
$C_5H_9NO_4(s)$	8.027×10^{22}	0.1333	19.62

15. 3.49×10^{-4} g

16. $CH_3COOH(\ell)$, $HOOCCOOH(s)$, $C_6H_5COOH(s)$

Answers to Chapter 5 Review Questions

1. c

2. a

3. b

4. a

5. d

6. c

7. b

8. c

10. 65.41; 6.02×10^{23}

14. Avogadro constant, 6.02×10^{23}

17. **a.** 0.36 mol
b. 15.8 mol
c. 5.91×10^{-22} mol

18. **a.** 1.9 g
b. 119 g
c. 1.4×10^5 g

19. 1.94×10^{-3} mol

20. 1.2×10^4 g

21. 20.0 g glucose = 0.111 mol; 20.0 g propane = 0.453 mol

22. NCl_3

23. a. 0.06802 mol
b. 0.08752 mol of octane
c. 0.09331 mol of cysteine

24.

Substance	Total number of atoms	Number of molecules or formula units	Molar mass (g/ mol)	Amount of substance (mol)	Mass (g)
$C_3H_6O_2(\ell)$	4.47×10^{23}	4.06×10^{22}	74.09	0.0675	5.00
$NaC_6H_5COO(s)$	3.56×10^{19}	2.37×10^{18}	144.11	3.94×10^{-6}	5.68×10^{-4}
$Al(H_2PO_4)_3(s)$	1.363×10^{24}	6.193×10^{22}	317.95	0.1029	32.71
$CCl_2F_2(g)$	2.38×10^{26}	4.75×10^{25}	121.91	0.0127	1.53
$C_4H_{10}O_2(\ell)$	5.23×10^{23}	3.46×10^{22}	90.14	0.0574	5.17
$NaHCO_3(s)$	1.778×10^{24}	2.963×10^{23}	84.01	0.4921	41.34

35. Ti_2S_3, titanium(III) sulfide

36. a. 1.2×10^{-5} mol

37. a. 2×10^{-6} mol

38. a. 0.42 g
b. 1.8×10^{-3} mol

39. a. 13 g of potassium nitrate, 0.28 g sodium fluoride
b. 0.12 mol potassium nitrate, 6.5 × 10—3 mol sodium fluoride

Answers to Chapter 5 Self-Assessment Questions

1. e
2. d
3. b
4. c
5. c
6. d
7. a
8. c
9. d
10. b

12. 0.52 g

13. 5.27×10^{-4} mol of Zn

14. 3.80×10^{2} g

15. b. 192.1 g/mol

16. 1.09×10^{21} molecules

17. 1.69×10^{22} carbon atoms; 9.87×10^{21} hydrogen atoms; 4.23×10^{21} chlorine atoms; 2.82×10^{21} oxygen atoms

18. a. 111.11 g/mol
b. 1.20×10^{25} carbon atoms; 1.20×10^{25} hydrogen atoms; 4.82×10^{24} oxygen atoms; 2.41×10^{24} nitrogen atoms

21. a. 4.25 kg
b. 14.3 kg

23. 57.51 g ethanol, 22.49 g water

25. 119 pg/day

Chapter 6

Answers to Caption Questions:

Figure 6.2: H_2O; two hydrogen atoms and one oxygen atom

Figure 6.4: The percent composition provides the basic information for determining the molecular formula of citric acid. From the formula a plan for synthesizing the compound can be created.

Figure 6.10: The numbers of atoms in the molecular formula are *n* times as great as the corresponding numbers in the empirical formula.

Answers to Learning Check Questions

1. Yes, water is the same regardless of the source, based on the law of definite proportions.

2. Sample answer: The elements can combine in different proportions to create different compounds with very different properties. For example: N_2O, NO_2, N_2O_5

3. The law is also called the law of definite composition or the law of constant composition because the chemical composition, including the ratio of elements in a compound, remains the same (that is, it is constant and definite).

4. The mass percent of carbon in carbon dioxide cannot change, based on the law of definite composition.

5. Carbon can be found in many different compounds with different formulas, so the mass percent can be different from one compound to the next.

6. a. 57.1 %.
b. Carbon and oxygen have different mass percents because their molar masses are different.

7. There is one atom of carbon for every four atoms of hydrogen; that is, the ratio is 1:4.

8. Both provide proportions of elements found in the compound; however, the proportions in percentage composition are based on the overall mass of each element found in the compound, and the proportions in a molecular formula are based on the number of atoms of each element in the compound.

9. The molecular formula shows the actual number of atoms of each element in a molecule of the substance, whereas the empirical formula only tells you the ratios of the atoms in a molecule.

10. The molecular and empirical formulas are the same when the actual amounts of each element inside a compound are already in the lowest whole number ratio.

11. Every compound has its own set of unique properties that is a direct result of its structure and composition; a molecular formula reflects the specific composition of a compound, whereas an empirical formula can represent two or more compounds with the same lowest whole number ratio. For example, NO_2 and N_2O_4, have the same empirical formula (NO_2) but have different properties.

12. NO_2 and N_2O_4; N_2O_4 is a whole-number multiple of NO_2.

13. A molecular formula includes the actual numbers of each atom in one molecule of a substance; in an empirical formula, although the numbers of each atom are in the correct ratios, they may not be the actual numbers that occur in one molecule of substance.

14. Choose a standard mass for the substance, such as 100 g. Then use the given mass percents to calculate the amount of mass of each atom in the 100 g sample. Then use the molar mass of each atom to determine the number of moles of each atom. Finally, divide or multiply each number of moles by the correct factor to convert each number to the smallest whole numbers possible.

15. The molar mass is determined experimentally, usually using a mass spectrometer.

16. A hydrate has a number of H_2O units attached to its molecule; an anhydrate does not.

17. Water molecules add mass to a solid and this extra mass can affect measurements and calculations.

18. Heat the substance to drive off all water molecules in the hydrate. Once all the water has been driven off, calculate the difference between the initial and final mass.

Answers to Practice Problem Questions

1. 22.27%
2. 69.55%
3. 7.8%
4. 53.28%
5. 25.6%
6. 27%
7. $H_2Cr_2O_7(aq)$
8. $H_2SO_3(aq)$
9. 63.89% Cl
10. $ZnS(s)$, $Cu_2S(s)$, $PbS(s)$
11. 82% N; 18% H
12. 68.4% Cr; 31.6% O
13. 40.0% C; 6.7% H; 53.3% O
14. 48% Ni; 17% P; 35% O
15. 37.0% C; 2.20% H; 18.5% N; 42.3% O
16. 67.10% Zn; 32.90% S
17. 127.8 g Cu; 32.2 g S
18. 24.74% K; 34.76% Mn; 40.50% O
19. 10.1% C; 0.80% H; 89.1% Cl
20. No, the percentage composition of carbon in the sample is 64.8%. If the sample was were ethanol, the percentage composition of carbon would be 52.1%.
21. 63.14% Mn, 36.86% S
22. 93.10% Ag, 6.90% O
23. 2.06% H, 32.69% S, 65.25% O
24. 34.59% Al, 61.53% O, 3.88% H
25. 41.40% Sr, 13.24% N, 45.36% O
26. 73.27% C, 3.85% H, 10.68% N, 12.20% O
27. 205 kg
28. 127 kg
29. 17.1 g
30. 248 kg
31. CH_3
32. Mg_2Cl
33. $CuSO_4$
34. $K_2Cr_2O_7$
35. NH_3
36. Li_2O
37. BF_3
38. Cl_3S_5
39. Na_2CO_3
40. P_2O_5
41. $C_6H_{12}O_6(s)$
42. $C_8H_{10}(\ell)$
43. $C_4O_2H_{10}(\ell)$
44. $C_8H_8(s)$
45. $HgCl(s)$
46. $C_8H_{10}N_4O_2(s)$
47. $C_2H_5NO_2(s)$
48. $B_2H_6(g)$
49. C_3ONH_8; $C_6O_2N_2H_{16}$
50. Its empirical formula is $C_9H_{12}O$ and its molecular formula is $C_{18}H_{24}O_2$
51. 50.88%
52. 48.08%
53. 13.43%
54. 62.97%
55. $MgSO_4 \cdot 7H_2O(s)$, $Ba(OH)_2 \cdot 8H_2O(s)$, $CaCl_2 \cdot 2H_2O(s)$, $Ca(SO_4)_2 \cdot 2H_2O(s)$, $Mn(SO_4)_2 \cdot 2H_2O(s)$
56. 5
57. 4
58. $Cr(NO_3)_3 \cdot 9H_2O$
59. $MgI_2 \cdot 8H_2O$
60. 2.83 g

Answers to Section 6.1 Review Questions

4. 42.10% C, 6.49% H, 51.41% O
6. 36% Ca, 64% Cl
7. 0.32% H, 57.95% Au, 41.73% Cl
10. 27.74% Mg, 23.57% P, 48.69% O
11. 7.52 g
12. a. 40.04% Ca, 12.00% C, 47.96% O
13. C
14. Na^+

Answers to Section 6.2 Review Questions

3. SnO_2
4. $AlCl_3$
5. $KMnO_4$
6. As_2O_3
7. $PbCl_2$
9. 89.16%
10. 7
11. 6
13. a. CF_2
  b. twice as much
14. $C_{10}H_{16}N_2O_8$
15. C_3H_6O
16. $C_{24}H_{24}O_6$

Answers to Chapter 6 Review Questions

1. b
2. d
3. a
4. d
5. c
6. e
7. a
8. b
9. b

10. a. 1:1; 2 H : 2 O
  b. 1:2:1; 2 C : 4 H : 2 O
  c. 3:1:4; 3 Na : 1 P : 4 O
  d. 1:1:3; 1 Ag : 1 N : 3 O

19. Mg_2Cl

20. $HBrO_3$

21. $C_6H_{12}O_6$

22. $C_4H_6O_6$

23. $C_{20}H_{40}\ O_2$

24. $C_{12}H_{12}Cl_9F_6$

25. $C_6H_6O_2$

26. CH_2

27. 43.09%

28. 80.48 g/mol

29. barium chloride dihydrate $BaCl_2.2H_2O$

30. hematite, Fe_2O_3 (s)

31. 49.47% C, 5.20% H, 28.85% N, 16.48% O

32. a. $Ba(OH)_2.8H_2O$; 45.6%
  b. $Na_2CO_3.10H_2O$; 62.9%
  c. $CoCl_2.6H_2O$; 45.24%
  d. $FePO_4.4H_2O$; 32.3%
  e. $CaCl_2.2H_2O$; 24.5%

33. $C_4H_4O_2$

39. Fe_3O_4

40. a. trial 1: 80.4% Zn(s), 19.6% O(g) ;
  trial 2: 80.4% Zn(s), 19.6% O(g) ;
  trial 3: 86.2% Zn(s), 13.8% O(g) ;
  trial 4: 80.4% Zn(s), 19.6% O(g)
  b. discard trial 3
  d. $x = 1, y = 1$

41. a. 64.67% NiS, 43.93% NiAs, 41.03% $(Ni,Fe)_9S_8$

42. a. 10.35%, 89.65%

42. a. mass percent of methane = 10.4%; mass percent of water = 89.6%

42. a. 10.4% methane; 89.6% water

43. 11.50 g, 12.99 g, 15.99 g, 17.48 g, 18.98 g, 20.48 g, 26.46 g

Answers to Chapter 6 Practice Test Questions

1. a
2. d
3. e
4. d
5. d
6. b
7. e
8. b
9. d
10. d

14. $BaCO_3$

15. H_2SO_3

16. N_2O_3

17. $NaNO_2$

18. 7

19. C_8H_{20}

20. N_2O_2

21. 68.9%

22. a. $SiCl_3$
  b. Si_2Cl_6

23. a. NO_2
  b. N_2O_4

24. a. Sodium carbonate heptahydrate
  b. 54.34%
  c. 52 kg

25. Cu_2S

Chapter 7

Answers to Caption Questions

Figure 7.1 8 slices of toast, 8 turkey slices, 4 lettuce leaves, and 4 tomato slices

Figure 7.3 20 atoms of H, 10 atoms of O

Answers to Learning Check Questions

1. 10 slices of toast, 10 turkey slices, 5 lettuce leaves, and 5 tomato slices

2. The exact proportion of moles of each reactant and product is needed to determine the relative amounts of reactants and products in a complete reaction. The coefficients show the reacting molecular and mole ratios.

3. One mole of methane reacts completely with two moles of oxygen to produce one mole of carbon dioxide and two moles of water.

4. the relative number of moles of each reactant and product in a complete reaction

5. Because we are interested in the mole ratio for each reactant and product, not for each individual atom. Coefficients refer to relative amounts of each entire molecule, whereas subscripts refer to the relative amounts of each atom within a molecule.

6. a. 2 mol $C_2H_6(g)$: 4 mol $CO_2(g)$ or 1:2
  b. 2 mol $C_2H_6(g)$: 7 mol $O_2(g)$ or 2:7
  c. 4 mol $CO_2(g)$: 6 mol $H_2O(g)$ or 2:3

7. The exact molar amount of a reactant or product, as predicted by a balanced chemical equation.

8. a. limiting reagent, gas; excess, oxygen in the air
  b. limiting reagent, deposits ($CaCO_3(s)$); excess, vinegar
  c. limiting reagent potato excess oxygen in the air

9. tomato

10. not necessarily; the limiting reactant is the one that is less than the stoichiometric amount (i.e., the reactant that would be used up while the other reactants are still available)

11. The amount in excess is not used in the reaction.

12. oxygen, because one expects there will be plenty of oxygen remaining after all the phosphorus has reacted with oxygen

13. The theoretical yield is the mass or amount of product calculated using the chemical equation and the associated reacting mole ratios. The actual yield is the mass or amount of product measured experimentally.

14. The theoretical yield is usually higher since some product is usually lost during the experiment no matter how carefully the experiment is done.

15. Improper lab techniques may reduce reaction yields in a number of ways. Some product may cling to various lab equipment and not be properly rinsed and collected. Spillage might occur. Measurements might not be made correctly.

17. If a reactant is not pure, then the actual mass of reactant that reacted is less than it should be; this will result in less product.

18. For example, a leak in a tank will cause the amount of fluid in the tank to decrease; a competing reaction drains away some of the desired product just like the leak in the tank.

If some of the turkey sandwiches are eaten there will be fewer on the plate to serve for lunch.

Possible answer: Children are eating cookies as they come out of the oven, reducing the actual yield.

Answers to Practice Problems Questions

1. 2 mol Mg(s) : 1 mol O_2(g) : 2 mol MgO(s)
2. 2 mol NO(g) : 1 mol O_2(g) : 2mol NO_2(g)
3. 1 mol Ca(s) atom : 2mol $H_2O(\ell)$: 1 mol $Ca(OH)_2$(s) : 1 mol H_2(g)
4. 2 mol C_2H_6(g) : 7 mol O_2(g) : 4 mol CO_2(g): 6 mol H_2O(g)
5. 5 molecules
6. 150 molecules of $AlCl_3$(s)
7. 3.4×10^{23}
8. 3.9×10^{24}
9. 2.72×10^{24}
10. 1
11. 0.25 mol
12. 6.00 mol
13. 4.50×10^4 mol
14. 3.6 mol O_2(g)
15. 4.70 mol
16. **a.** 46.8 mol
  b. 187 mol
17. 56.5 mol O_2(g)
18. 6.45 mol of P_4(s)
19. 5.1×10^{23}
20. 7.24×10^5 mol of H_2O(g)
21. 1.3×10^5 g
22. 860 kg
23. 726 g
24. 0.123 mol
25. 0.421 g
26. 4.35 g
27. 10.3 g
28. 0.963 g
29. 5.16×10^{-2} g
30. 21.0 g
31. CaF_2(s)
32. $C_7H_6O_3$(aq)
33. $H_2O(\ell)$
34. $NiCl_2$(aq)
35. HNO_3(aq)
36. Li(s) is limiting, O_2(g) is in excess
37. $Na_2S_2O_3$(aq)
38. C_3H_6(g)
39. 61 g
40. **a.** oxygen
  b. 8
41. **a.** $AgNO_3$(aq)
  b. 0.03519 moles
42. 66.98 g
43. 11.4 g
44. 172 g
45. 9.8 g
46. 57.3 g
47. 11.8 g
48. 69.94 g
49. **a.** 0.446 g
  b. F_2(g), 24.0 g
50. **a.** 388.4 g of $CaCl_2$(aq) is excess (the total initial mass of $CaCl_2$(aq) is 776.9 g)
  b. 1189 g of $AgNO_3$(aq)
  c. 1003 g of AgCl(s) and 574.3 g of $Ca(NO_3)_2$(aq)
51. 0.97 g, 0.77 g
52. **a.** 37.3 g
  b. 75.6 %
53. 79.3%
54. **a.** 8.32 g
  b. 5.99 g
55. 44.6 g
56. 11.9 g
57. 61.8 g
58. 0.46 g
59. **a.** 1.51 kg
  b. 1.18 kg
60. 35 g

Answers to Section 7.1 Review Questions

4. 6.0×10^2 g
5. 28 g
7. **a.** $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$
  b. 526.2 kg or about ½ tonne
9. 2, 1, 1, 2; 1.20×10^{24}, 6.02×10^{23}, 6.02×10^{23}, 1.20×10^{24}; 162 g, 74.0 g, 200 g, 36.0 g
10. **a.** 1.7×10^{22}
  b. 2.30×10^{22} formula units

Answers to Section 7.2 Review Questions

1. water
2. O_2(g)
3. O_2(g)
4. **a.** $FeCl_3$(aq)
  b. 2.83 g NaOH(aq) would remain
  c. 37.8 g of $Fe(OH)_3$(s), 62.0 g of NaCl(s)
5. 3.47 g

9. 15.75 g

10. b. 0.15 mol

Answers to Section 7.3 Review Questions

4. 94.60%

5. 91.9%

11. a. 3.4 g of I_2, 1.6 g of NaCl
  b. 1.0 g of NaCl

12. Theoretical Yield: 2.638 g $CuSO_4(s)$, 1.4890 g H_2O; Actual Yield: 2.913 g $CuSO_4(s)$, 1.214 g H_2O; Percentage Yield: 110.4% $CuSO_4(s)$, 81.5% H_2O.

Answers to Chapter 7 Review Questions Questions

1. b
2. d
3. c
4. b
5. d
6. a
7. a
8. e

17. a. $2Al(s) + 3Br_2(g) \longrightarrow 2AlBr_3(s)$
  b. 7.5 mol
  c. 5.0 mol

18. theoretical yield: 14.77 g; percent yield: 98.75%

19. 2.3 mol

20. 0.488 g, 93.2%

22. b. 3.86 g

24. 87.3%

Answers to Chapter 7 Self-Assessment Questions

1. b
2. b
3. a
4. e
5. c
6. a
7. e
8. b
9. a

12. 17.81 g

13. a. 4.68×10^{-2} mol
  b. 4.5 mol

15. 93.2 kg

16. a. $3Zn(s) + 2FeCl_3(aq) \longrightarrow 2Fe(s) + 3ZnCl_2(aq)$
  b. 272.73 g
  c. 479.12 g

19. b. 55.78 %

20. 0.31 g

22. 0.185 mol

23. O_2 limiting reactant

25. 12.5 mol

Answers to Unit 3 Review Questions

1. e
2. e
3. c
4. a
5. b
6. e
7. b
8. b

12. a. $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
  b. 1.07 mol; 2.16 g

14. 19 kg

15. a. $O_2(g)$
  b. 19.1 g

16. 17.2 g

17. 1.97×10^{23} g
  b. 1.66×10^{-22} mol per meal, 32.8 g per meal
  c. 2×10^{-11} mol
  d. 4.13×10^{12} g

18. 27 % K, 35 % Cr, and 38% O

19. 32.37% Na, 22.58% S, 45.05% O

20. a. $Pb_1S_2O_8$; $Pb(SO_4)_2$
  b. lead(IV) sulfate

21. Sb_2S_3

22. $Na_2C_8H_4O_4$

23. $C_3H_5O_2I$; $C_6H_{10}O_4I_2$

24. a. 40% C, 6.7% H, 53.3% O; CH_2O
  b. $C_5H_{10}O_5$

27. 0.3 g Al

32. a. $2C_8H_{18}(\ell) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$
  b. 0.700 mol; 30.8 g

35. HCO_2

37. C_3H_6O

38. a. 115 kg
  b. 144 kg
  c. 228 kg
  d. 81.6 %

39. $Na_2SO_4 \cdot 10H_2O$, sodium sulfate decahydrate

40. 11.99 g Ag/1.0 g Al

41. $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$; $2C_8H_{18}(\ell) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$

42. a. Two cartridges are needed; 15.1 kg H_2O
  b. 51.7 kg $CO_2(g)$

Answers to Unit 3 Self-Assessment Questions

1. d
2. c
3. b
4. a
5. b
6. d
7. c
8. e
9. b
10. c

11. 11.2 g

13. CH_2; C_6H_{12}

15. Cu_2S

16. $C_{20}H_{12}O_4$

17. 16.0 g

21. $C_{14}H_{18}N_2O_5$

23. 26.7%

Unit 4

Chapter 8

Answers to Learning Check Questions

1. The mixture is not homogeneous on the microscopic scale.

2. Nitrogen is the solvent, because it is present in the largest proportion.

3. Sample answers: solid solute: sugar in water; liquid solute: alcohol in water; gas solute: carbon dioxide in water.

4. Gasoline and water do not dissolve in one another (they are immiscible).

5. Yes. A solute is classified as insoluble if less than 0.1 g will dissolve in 100 mL of solvent. Therefore, if as much solute is dissolved in the solvent as possible, the solution will be dilute but saturated.

6. Yes. When the pressure is released by opening the can, the excess solute (carbon dioxide gas) leaves the solution.

7. No. Like dissolves like, so if benzene is a solvent for fats and oils, which are non-polar, it should not dissolve in a polar solvent such as water.
8. A fluoride ion has a very small radius. This results in most fluorides being less soluble than the other halides.
9. Methanol. The solubility of liquids is changed very little by temperature.
10. The taste is affected by how much carbon dioxide is dissolved in the soft drink. The solubility of carbon dioxide in water decreases at higher temperature.
11. Unsaturated. At 40°C, the solubility of $KNO_3(s)$ is about 53 g/100 g.
12. Fish avoid extremes of hot and cold, just as we do. In particular, warm water contains less dissolved oxygen, so fish are less likely to be found in these areas.
13. The concentration is likely to be more precise.
14. A volumetric flask and a volumetric pipette, or graduated pipette, can be used to prepare a solution with a known concentration.
15. 1. Measure the mass of pure solute and dissolve in water. Transfer to a 1 L volumetric flask and make up the solution to the calibration mark. 2. Dilute a more concentrated standard solution using a graduated pipette and a 1 L volumetric flask.
16. $n = c \times V$; the product of concentration and volume is constant. Thus, as the volume of solution increases, its concentration decreases.
17. Drying the beaker is not necessary in this case. The water in the beaker will dilute the solution slightly, but it will not change the amount of $CuSO_4(aq)$, which is determined by the volume of the pipette and the concentration of the solution in it.
18. First prepare a more concentrated solution and dilute it to the necessary concentration.

Answers to Practice Problems

1. 1.2×10^2% (m/v)
2. 8.6% (m/v)
3. 31.5 g
4. 8.0×10^1 g
5. 6.67% (m/v)
6. 1.75% (m/v)
7. 131 g
8. 3 g NaCl, 0.09 g KCl, 0.1 g $CaCl_2$
9. 800 mL
10. Measure 14 g of the solute and dissolve in water. Add water to bring the total volume of the solution to 400 mL.
11. 15% (m/m)
12. 3.8% (m/m)
13. 90 kg chromium, 40 kg nickel, 370 kg iron
14. 7.91% (m/m)
15. 15.1% (m/m)
16. 11 g
17. 5 g
18. 1.7% (m/m)
19. 7.23×10^{-4}% (m/m)
20. 0.243% (m/m) nickel, 2.3% (m/m) copper, 2.3×10^{-4}% (m/m) platinum
21. 400 mL
22. 20% (v/v)
23. 2 L
24. about 2.2 L
25. Add enough water to increase the volume to 1.00 L
26. 2 L
27. 9% (v/v)
28. 6 L
29. 4.0×10^1% (v/v), assuming that the mixed solution has a volume of 20.0 L
30. 74 mL
31. 7.2 ppm
32. 35 mg
33. 0.7 ppm
34. 2×10^{-6} g
35. 11.5 ppm
36. The concentration is 3.0 mg/L, within limits.
37. 0.1 g or 100 mg
38. 4 g
39. 0.52 μg or 5.2×10^{-7} g
40. 1×10^2 ppm or 100 ppm
41. **a.** 1.5 mol/L
    b. 0.2 mol/L
42. 1.4 L
43. **a.** 1 mol/L
    b. 0.628 mol/L
    c. 0.1 mol/L
44. 0.14 g
45. **a.** 0.093 g
    b. 10 g
    c. 534 g
46. **a.** $Na^+(aq) = 1.2$ mol/L; $SO_4^{2-}(aq) = 0.60$ mol/L
    b. $NH_4^+(aq) = 3.1$ mol/L; $PO_4^{3-}(aq) = 1.0$ mol/L
    c. $Ca^{2+}(aq) = 1 \times 10^{-4}$ mol/L; $PO_4^{3-}(aq) = 8 \times 10^{-5}$ mol/L
47. 0.29 mol/L
48. 12 mol/L
49. 218 mL
50. 2.0×10^{-5} mol/L
51. **a.** 33.3 mL
    b. 107 mL
    c. 25 mL
52. **a.** 0.99 mol/L
    b. 0.19 mol/L
    c. 0.0116 mol/L
53. 0.0938 L
54. 0.02 mol/L
55. **a.** 0.08 mol/L
    b. 0.2 mol/L

56. 15% (m/v)

57. 3.00×10^2 mL

58. 0.5 L; about 0.5 L

59. All parts: Your procedure should be similar to the procedure outlined in **Table 8.7**.

59. a. Mass 2.1 g $AgNO_3(s)$
  b. Mass 6.05 g $K_2CO_3(s)$
  c. Mass 12.6 g $KMnO_4(s)$

60. All parts: Your procedure should be similar to the procedure outlined in **Table 8.8**.
  a. Dilute 29 mL
  b. Dilute 7.5 mL
  c. Dilute 945 mL

Answers to Section 8.3 Review Questions

3. a. 0.04% (m/m)
  b. 4×10^2 ppm

4. 0.5 g

6. 23.9% (v/v)

7. 8.7 mol/L

8. K_3PO_4 or $K_3C_6H_5O_7$

9. 5.3×10^{14} kg

10. 0.7 mol/L

11. 0.098 mg

12. 45 ppb

13. No. The ground water concentration is over 4000 ppb.

14. 0.4 mol/L

Answers to Caption Questions

Figure 8.8 Sucrose is not an ionic compound.

Figure 8.11 The solubility is about 45 g KCl(aq) per 100 g water.

Figure 8.17 Iron is present in the greatest amount in the solution.

Figure 8.19 Sample answer: A pair of glasses has a frame and two lenses.

Figure 8.20 All of the flasks have long, thin necks, stoppers, and a graduation mark to indicate volume.

Answers to Section 8.4 Review Questions

3. 0.159 mol/L

4. 33 mL of stock solution is required

5. 3.00 L

6. 250 mL; 500 mL

7. 4.10 g

8. 7.5×10^9 L

11. 0.50 L of stock solution is required

12. 567 mL

14. 5.105 g of KHP is required

Answers to Chapter 8 Review Questions

1. c
2. b
3. d
4. d
5. e
6. e
7. d
8. c

18. a. 45 mL
  b. 64 mL

19. 3.9% (m/m) zinc, 1% (m/m) tin, 95% (m/m) copper

20. 4×10^{-5} mol

21. 55.49 mol/L

22. a. 0.204% (m/v)
  b. 8.60×10^2 ppm
  c. 0.14 mol/L
  d. 1.0×10^{-2} mol/L
  e. 4.5×10^{-2} mol/L

23. 2.2 L

24. b. approximately 90 g
  c. approximately 76°C

25. a. approximately 380 g
  b. 40°C

35. a. 0.1 ppb
  b. 5×10^{-10} mol/L

Answers to Chapter 8 Self-Assessment Questions

1. c
2. b
3. e
4. c
5. b
6. e
7. c
8. b
9. c
10. b

17. 85 g/100 mL

18. 50% (m/m)

19. 29 mL

20. 0.4 ppm

21. 200 L

22. 943 mL; about 2.06 L of distilled water

24. use 4.0 g of $KMnO_4(s)$

25. a. $AgNO_3(aq)$

26. 205 mL

Chapter 9

Answers to Learning Check Questions

1. Any insoluble substance, such as AgCl(s), $CaCO_3(s)$, or a molecular compound such as $H_2O(\ell)$ or $CO_2(g)$ is never shown as an ion in a net ionic equation.

2. A chemical equation shows reacting substances; an ionic equation shows soluble ionic substances in their dissociated form.

3. Atoms and electrons cannot be gained or lost during a chemical change.

4. There is no reaction because all the ions are soluble.

5. Flame test: the colour of the flame when a substance is burned can indicate what elements are present; solution colour: the colour of a solution can indicate what ions are present in the solution; precipitate: if a precipitate forms under specific reaction conditions, it is possible to deduce the ions that were present in the solution.

6. For most substances, there is no relationship between the colour of a solution, which is usually (but not always) due to certain anions, and the colour in a flame test, which is due to certain metal cations.

7. n is amount of a given substance (in moles), and V is volume of solution (in litres)

8. You need to know the molar ratios of the reactants and products.

9. The molar mass of the solute is needed.

10. The limiting reactant is the reactant that is completely consumed during a chemical reaction, limiting the amount of product that is produced

11. The cations and anions have equal but opposite charges.

12. The calculations are based on molar ratios and amounts in moles.

13. If your water is hard, you will have difficulty in producing a lather using soap, or you might observe lime build-up on a heating element. If your water is soft, you will find that soap lathers easily, and you should not have notable build-up on heating elements, dishes, or shower tiles.

14. $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ are present in greater concentrations in hard water.

15. The amount of calcium carbonate and magnesium carbonate in rocks varies in different areas; these ions dissolve in rainwater as it flows through the rocks.

16. Arsenic-containing soil was deposited as silt in a river delta. Deep wells penetrate below the layer of silt.

17. The ions $AsO_4^{3-}(aq)$ and $PO_4^{3-}(aq)$ have the same charge and are similar in radius. Thus, they should have similar solubility when combined with a cation.

18. The signs of dental fluorosis are streaking in the teeth, and especially brown mottling. You should suggest that the patient get his/her water tested for $F^-(aq)$, if using well water. You could also suggest using a fluoride-free toothpaste and mouth wash.

19. Temporary hardness is the result of dissolved $CaCO_3(aq)$ and $MgCO_3(aq)$ and can be removed by boiling water. Permanent hardness results from dissolved $CaSO_4(aq)$ and $MgSO_4(aq)$ and can only be removed by chemical means.

20. The salt is used to regenerate sodium ions on the resin beads inside the unit, after the beads have become coated with calcium or magnesium ions.

21. Oil can be supplied to both plants through the same pipeline. Also, waste heat from the electrical generating plant can supply some of the heat needed by the desalination plant.

22. Canada has sufficient supplies of fresh water.

23. There has been a steady increase in life expectancy since 1800, as water treatment techniques have improved.

24. Filtering, removing suspended particles, killing harmful bacteria.

Answers to Practice Problem Questions

1. a. $3Ba^{2+}(aq) + 2PO_4^{3-} \rightarrow Ba_3(PO_4)_2(s)$

2. $Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$

3. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

4. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$; $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

5. The precipitate is FeS(s). The spectator ions are $Na^+(aq)$ and $SO_4^{2-}(aq)$. The net ionic equation is: $Fe_2^+(aq) + S_2^-(aq) \rightarrow FeS(s)$

6. a. Spectator ions: $NH_4^+(aq)$ and $SO_4^{2-}(aq)$; net ionic equation: $3Zn_2^+(aq) + 2PO_4^{3-}(aq) \rightarrow Zn_3(PO_4)_2(s)$

b. Spectator ions: $Li^+(aq)$ and $NO_3^-(aq)$; net ionic equation: $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(\ell)$

c. No spectator ions; $2H^+(aq) + SO_4^{2-}(aq) + Ba^{2-}(aq) + 2OH^-(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$

7. $Pb_2^+(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

8. Examples: $Al(NO_3)_3(aq)$ or $Al(CH_3COOH)_3(aq)$ and $Na_2Cr_2O_7(aq)$ or $(NH_4)_2Cr_2O_7(aq)$

9. $Fe_3^+(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$; The spectator ions are $NO_3^-(aq)$ and $K^+(aq)$

10. a. $Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3(aq)$; $Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$

b. $Co(CH_3COO)_2(aq) + (NH_4)_2S(aq) \rightarrow CoS(s) + 2NH_4CH_3COO(aq)$; $Co^{2+}(aq) + S^{2-}(aq) \rightarrow CoS(s)$

11. $NH_4^+(aq) = 0.4$ mol/L; $PO_4^{3-}(aq) = 0.1$ mol/L

12. 0.11 mol/L

13. $2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$; 0.239 mol/L

14. 1.10 g FeS(s)

15. 2.7 g

16. 24 g

17. 1.22 g

18. 0.0370 mol/L

19. a. 0.214 mol/L

b. 12.5 g

20. a. 0.40 mol/L

b. 16 g

21. 12 g

22. 0.518 g

23. 96.4 mL

24. 0.662 g

25. 4.11 g; The calculation assumes $PbI_2(s)$ is completely insoluble, whereas some lead(II) iodide will remain in solution.

26. 33.5 mL

27. 0.826 mol/L

28. 19.5 mL

29. 0.500 g

30. 0.0150 mol/L

Answers to Caption Questions

Figure 9.4: Sample answer: red: Li^+ and/or Sr_2^+, green: Cu_2^+ and/or Ba_2^+, orange: Ca_2^+ and/or Na^+

Figure 9.6: $CuCl_2$ is soluble, while AgCl is not.

Figure 9.8: Over 99% of earth's water is unavailable for human consumption.

Figure 9.18: Benefits of a home water softener: reduced deposits on heating elements, better soap lather; Disadvantages: cost, maintenance, people on low-sodium diets cannot use a water softener

Figure 9.21: The 1900s had the greatest increase in life expectancy.

Answers to Section 9.1 Review Questions

2. a. $Cl^-(aq)$ and $NH_4^+(aq)$

b. $NO_3^-(aq)$ and $Ba^{2+}(aq)$

c. $Na^+(aq)$ and $Cl^-(aq)$

3. a. $3Cu^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Cu_3(PO_4)_2(s)$

b. $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

c. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

4. a. copper(II) carbonate, $CuCO_3(s)$
  b. $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
  c. $Na^{+}(aq)$ and $SO_4^{2-}(aq)$

5. a. Sample answers: $BaCl_2(aq)$, $BaBr_2(aq)$, $BaI_2(aq)$, or $Ba(OH)_2(aq)$ and $Na_3PO_4(aq)$, $K_3PO_4(aq)$ or $(NH_4)_3PO_4(aq)$
  b. Sample answers: $MgCl_2(aq)$, $MgBr_2(aq)$, $MgI_2(aq)$, $MgSO_4(aq)$ and $NaOH(aq)$, $KOH(aq)$
  c. Sample answers: $Al(NO_3)_3(aq)$ or $Al(CH_3COO)_3(aq)$ and $Na_2Cr_2O_7(aq)$ or $K_2Cr_2O_7(aq)$

11. a. $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$; $Ca^{2+}(aq) + 2OH^{-}(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$

12. a. $Pb^{2+}(aq)$, $PbI_2(s)$
  b. $Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$

14. a. $H_2SO_4(aq)$
  b. $HCl(aq)$

Answers to Section 9.2 Review Questions

1. 0.10 mol/L $MgCl_2(aq)$ There are 2 mol of chloride ions for every 1 mol of magnesium chloride.

2. a. 0.5 mol/L
  b. 0.45 mol/L

3. 42 mL

4. 4.00×10^{2} mL

5. 135 mL

6. a. 0.66 g $H_2(g)$
  b. 1.5 mol/L

7. a. $H_3C_6H_5O_7(aq)$ (citric acid)
  b. 0.238 g

8. $PbSO_4$; 2 g

9. 2×10^{-3} mol/L

10. 2.00×10^{2} mL $CaCl_2(aq)$ and 99.9 mL $K_2CO_3(aq)$

12. 0.64 mL

Answers to Section 9.4 Review Questions

16. a. Sample answers: $NaCl(aq)$; $Na_2SO_4(aq)$
  b. Sample answers: $Pb^{2+}(aq) + 2Cl^{-}(aq) \rightarrow PbCl_2(s)$; $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

Answers to Chapter 9 Review Questions

1. b
2. b
3. c
4. d
5. a
6. b
7. c
8. e

10. $Na^{+}(aq)$, $SO_4^{2-}(aq)$

11. a. $Cu(s) + Fe^{2+}(aq) \rightarrow Cu^{2+}(aq) + Fe(s)$
  b. $Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$
  c. $H^{+}(aq) + OH^{-}(aq) \rightarrow H_2O(\ell)$

12. a. $Cu^{2+}(aq)$
  b. $Ca^{2+}(aq)$
  c. $Na^{+}(aq)$ and $MnO_4^{-}(aq)$

17. 0.010 mol

18. 1.5 mol/L Mg^{2+}; 3.0 mol/L NO_3–; One formula unit of magnesium nitrate dissociates into one magnesium ion and two nitrate ions.

19. b. $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(\ell)$; $Mg(OH)_2(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(\ell)$

21. 0.015 mol/L

22. a. $Na^{+}(aq) = 5.91$ mol/L; $NO_3^{-}(aq) = 5.91$ mol/L
  b. $Ca^{2+}(aq) = 0.88$ mol/L; $CH_3COO^{-}(aq) = 1.8$ mol/L
  c. $NH_4^{+}(aq) = 1$ mol/L; $PO_4^{3-}(aq) = 0.4$ mol/L

23. 0.22 g $CaCO_3(s)$

24. $Zn(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Zn^{2+}(aq)$; 0.096 mol/L

25. $Cu^{2+}(aq) + 2F^{-}(aq) \rightarrow CuF_2(s)$; 1.46 mol/L

26. a. $MgCl_2(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + 2NaCl(aq)$
  b. $Mg^{2+}(aq) + 2OH^{-}(aq) \rightarrow Mg(OH)_2(s)$
  c. 0.9 g

35. 0.3 mol/L

36. b. $(NH_2)_2CO\ (aq) + H_2O(\ell) \rightarrow 2NH_3(aq) + CO_2(g)$; $2NH_3(aq) + 3O_2(g) \rightarrow 2NO_2^{-}(aq) + 2H^{+}(aq) + 2H_2O(\ell)$; $2NO_2^{-}(aq) + O_2(g) \rightarrow 2NO_3^{-}(aq)$

39. a. 4 L

Answers to Chapter 9 Self-Assessment Questions

1. c
2. e
3. a
4. d
5. b
7. a
8. b
9. e
10. d

12. a. $2Na_3PO_4(aq) + 3Ca(OH)_2(aq) \rightarrow 6NaOH(aq) + Ca_3(PO_4)_2(s)$
  b. $3Ca^{2+}(aq) + 2PO_4^{3-}\ (aq) \rightarrow Ca_3(PO_4)_2(s)$

13. a. tin(II) phosphate, $Sn_3(PO_4)_2(s)$
  b. $K^{+}(aq)$ and $Cl^{-}\ (aq)$
  c. $3Sn^{2+}(aq) + 2PO_4^{3-}\ (aq) \rightarrow Sn_3(PO_4)_2(s)$

14. $Cr^{2+}(aq)$, $Cu^{2+}(aq)$; A bluish flame in a flame test would show that the solution contains copper(II) instead of chromium(II). Also, hydrogen sulfide could be added: copper(II) will form a precipitate with the sulfide ions.

15. Sample answers: Use $NaCl(aq)$: $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$; Use $Na_2SO_4(aq)$: $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$

16. 5.6 g

17. a. 2.0 mol
  b. 3×10^{3} kg

18. 25 g

Chapter 10

Answers to Learning Check Questions

1. An Arrhenius acid is a substance that ionizes in water to produce one or more hydrogen ions, $H^{+}(aq)$. An Arrhenius base is a substance that dissociates in water to form one or more hydroxide ions, $OH^{-}(aq)$.

2. a. Sample answers: citrus fruit, tomatoes, vinegar, carbonated drink
  b. Sample answers: soap, detergent, ammonia solution (window cleaner), oven cleaner

3. Sample answer: Venn diagram: two circles that partially overlap, one labelled “Acids”, one labelled “Bases”. Within the overlapping section: conduct electricity and corrode tissues. In the “Acids” circle: sour taste, pH less than 7, turn litmus red, produce no colour change in phenolphthalein, corrode metals, react with active metals to produce hydrogen gas, react with carbonates to produce carbon dioxide gas. In the “Bases” circle: bitter taste, slippery texture, pH greater than 7, turn litmus blue, turn phenolphthalein pink, do not corrode metals, no reaction with active metals, no reaction with carbonates.

4. a. HBr(aq): acid
  b. KOH(aq): base
  c. H_3PO_4(aq): acid
  d. $HClO_4$(aq): acid
  e. $Ca(OH)_2$(aq): base
  f. HNO_3(aq): acid
  g. $Sr(OH)_2$(aq): base
  h. CsOH(aq): base

5. The paper will remain blue, because the solution is basic.

6. a. The H^+(aq) concentration is higher than the OH^-(aq) concentration.
  b. The H^+(aq) concentration is higher than the OH^-(aq) concentration.
  c. The OH^-(aq) concentration is higher than the H^+(aq) concentration.

7. A strong acid is one that ionizes completely into ions; for example, HCl(aq). A concentrated acid is one with a relatively large amount of acid dissolved in the solution; for example, 6 mol/L HCl(aq).

8. a. 0.01 mol/L NaOH(aq)
  b. 4 mol/L HF(aq)

9. Both acids are weak. Reasoning: Soft drinks are consumed by humans and it is dangerous and deadly to consume strong acids. Also, neither is on the list of strong acids.

10. The acids are listed in order of increasing strength. The addition of more oxygen atoms increases the polarity of the bond between the ionizable hydrogen atom and the oxygen atom it is attached to.

11. Your diagrams should be similar to **Figure 10.9** on page 461. They should show a high degree of dissociation in a strong base, compared with a high concentration of solute in a concentrated base, and a low degree of dissociation in a weak base, compared with a low concentration of solute in a dilute base.

12. The safety hazards associated with strong acids and bases are far greater than those associated with weak acids and bases. Strong acids and bases are highly corrosive and should never be consumed, while weak acids and bases are actually ingredients in some common foods and beverages.

13. A salt and water are produced by a neutralization reaction.

14. The pH at the equivalence point is usually not 7.0 for titrations involving either a weak base or a weak acid.

15. If both concentrations were unknown, you would have no way to calculate the concentration of either the base or the acid. The volume and concentration of the known solution and the volume of the unknown solution at the equivalence point are all needed to calculate the unknown concentration.

16. The term “neutralization reaction” refers to a reaction between an acid and a base. It does not mean that all of the hydrogen ions and all of the hydroxide ions have been neutralized.

17. The end point occurs when the indicator changes colour. The equivalence point occurs when stoichiometric quantities of acid and base have been mixed together.

18. Phenolphthalein changes colour in basic solution between pH 8.2 and pH 10.0. Thus, more of the weak base will be added than required for equivalence.

Answers to Practice Problem Questions

1. 2.12 mol/L
2. 67.5 mL
3. 0.1298 mol/L
4. 107 mL
5. 0.32 mol/L
6. 0.128 mol/L
7. 2×10^{-4} mol/L
8. 87.3 mL
9. 118 mL
10. 2×10^1 mL (Rounded to appropriate number of significant digits)

Answers to Caption Question

Figure 10.1 Both lemons and grapes are acidic.

Answers to Section 10.1 Review Questions

8. a. pH < 7
  b. pH > 7

10. $H_2O(\ell)$, H^+(aq), and Cl^-(aq) are present in an aqueous solution of hydrochloric acid.

13. NH_3(aq), CH_3COOH(aq), HCl(aq) Ammonia is a weak base, acetic acid is a weak acid, and hydrochloric acid is a strong acid.

Answers to Section 10.2 Review Questions

1. acid + base → salt + water

2. a. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(\ell)$
  b. $2HI(aq) + Mg(OH)_2(aq) \rightarrow MgI_2(aq) + 2H_2O(\ell)$

8. 0.5 mol

10. 21.3 mL

13. 0.1538 mol/L

Answers to Chapter 10 Review Questions

1. d
2. a
3. c
4. c
5. a
6. b
7. d
8. e

9. Sample answers: sodium hydroxide, NaOH(aq), potassium hydroxide, KOH(aq), ammonia, NH_3(aq), magnesium hydroxide, $Mg(OH)_2$(aq)

10. a. Sample answers: hydrochloric acid, $HCl(aq)$, hydroiodic acid, $HI(aq)$
  b. Sample answers: sulfuric acid, $H_2SO_4(aq)$, carbonic acid, $H_2CO_3(aq)$
  c. Sample answers: phosphoric acid, $H_3PO_4(aq)$, citric acid, $H_3C_6H_5O_7(aq)$

12. a. $Ca(OH)_2(aq) + 2HCl(aq) \rightarrow 2H_2O(\ell) + CaCl_2(s)$
  b. $2H_3PO_4(aq) + 3Sr(OH)_2(aq) \rightarrow 6H_2O(\ell) + Sr_3(PO_4)_2(s)$
  c. $2NaOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(\ell) + Na_2SO_4(s)$

18. $HOCl(aq) \rightarrow H^+(aq) + OCl^-(aq)$

20. 0.101 mol/L

23. 0.5 mol

24. a. 10.90 mL, 10.05 mL, 9.95 mL
  c. 0.21 mol/L

35. The pH of the solution is between 2.8 and 6.0.

37. a. 8×10^1 kg (Rounded to appropriate number of significant figures)
  b. 3.68 mol/L

39. 0.429 g

40. Volume of MOH neutralized should be 100.0 mL. Based on this revised information, the molar mass of MOH is 24.0 g/mol.

41. 67.6% (m/m)

Answers to Chapter 10 Self-Assessment Questions

1. b
2. a
3. b
4. e
5. e
6. b
7. c
8. b
9. b
10. d

22. The pH is less than 7.

24. 0.2479 mol/L

25. 0.585 mol/L NaOH(aq)

Answers to Unit 4 Review Questions

1. e
2. a
3. c
4. d
5. a
6. a
7. c
8. c

12. $K^+(aq)$ and Cl^- (aq)

19. You could add sulfuric acid, which would precipitate Ca^{2+}, but not Mg^{2+}. Equation: $Ca(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(s) + 2HNO_3(aq)$

20. 3.0 mol/L

21. Assuming the density of solution is 1.0 g/mL, the atomic mass of X is 19 g/mol and the anion is F^-.

22. Precipitate: iron(III) hydroxide, $Fe(OH)_3(s)$ Net ionic equation: $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

23. 0.137 g

24. a. 0.3 mol/L
  b. 0.7 mol/L

26. 0.716 mol/L

29. The pH is less than 3.2.

30. a. 3.67 g

31. a. 4 L

32. 6×10^2 mL

33. 0.886 mol/L

49. 210 mL

50. 0.05 ppm

55. a. 3 × 107 L

57. a. 0.7660 mol/L
  b. 4.02% (m/v)

Answers to Unit 4 Self-Assessment Questions

1. b
2. b
3. d
4. e
5. d
6. c
7. c
8. b
9. c
10. c

14. 20 mL

15. 4.5×10^{19} t You need the density of sea water; assuming it is 1 g/mL, then the answer is as stated.

18. a. $Ba(NO_3)_2(aq) + Na_2S(aq) \rightarrow 2NaNO_3(aq) + BaS(s)$
  b. 0.2 mol/L

25. 76.9 mL

Unit 5

Chapter 11

Answers to Learning Check Questions

1. The solid state of water is commonly called ice. The liquid state is always called water and the gaseous state is called water vapour or steam.

2. Stronger the attractive force, more likely to be a solid or liquid at room temperature

3. The molecules in gases are more random than in a liquid, and in solid the molecules are more ordered than in a liquid.

4. Intermolecular forces hold the molecules together in an orderly fasion; without them the molecules would be free to move randomly about. Without intermolecular forces all substances would be gases.

5. The average kinetic energy of the particles increases as you go from solid to liquid to gas.

6. Graphic organizers should distinguish between electrostatic attraction, interactions between polar molecules, and interactions between non-polar molecules.

7. It is the force that a column of air exerts on the Earth's surface divided by the area of Earth's surface at the base of the column.

8. The height of mercury in a tube immersed in a shallow dish of mercury was dependent on atmospheric pressure. The higher the atmospheric pressure, the higher the level of mercury in the tube.

9. a. 363 kPa
  b. 1.39 atm
  c. 1.0×10^2 kPa
  d. 6.3×10^4 Pa

10. number of bars = number of psi × (1.01325/14.7)

11. As altitude increases, atmospheric pressure decreases. This means there is less oxygen available at a higher altitude and breathing may need to be assisted.

12. The force on the surface of the water in the pan due to atmospheric pressure is balanced by the pressure from the water in the bottle, which tends to push the water in the pan upward. When the water level in the pan falls below the water level in the bottle, the force from the water in the bottle is greater than the force due to atmospheric pressure, pushing the water level in the pan back up until it is level with the water in the bottle.

13. For a given amount of gas at constant pressure the volume of a gas varies directly with the temperature expressed in K.

14. Absolute zero is the extrapolated value for the volume-temperature graph when the volume of the gas is zero—an impossible situation for real gases. Kelvin suggested that this temperature, −273.15°C, is the lowest possible temperature value. A new temperature scale was created with 0 K as the starting temperature.

15. Linear lines with common *x*-intercept at −273.15°C

16. **a.** 300 K
  b. 2.48×10^2 K
  c. 100.1°C
  d. -2.5×10^2°C

17. Volume of a gas is inversely related to pressure of gas (Boyle's Law) and it, therefore, represents another variable that will affect the study of the gas.

18. The balloon quickly shrivels upon contact with the liquid nitrogen as molecular motion slows down with temperature decrease; the volume of the gas decreases.

Answers to Practice Problems

1. 2.29 atm
2. 398 kPa
3. 24 atm
4. 0.27 L
5. 1.3×10^2 kPa
6. 440 L
7. 14.3 mL
8. 1.1×10^2 L
9. 1.73×10^3 L
10. **a.** 3.6×10^2 L
  b. 5.6×10^2 min
11. 11 L
12. 35.5 mL
13. 1.29 L
14. 95 L
15. 561 cm^3
16. 539 K
17. 308 K
18. 27°C
19. 1.25 times room temperature
20. −214°C
21. 42.0 kPa
22. 327 kPa
23. 2800 kPa
24. 430 kPa
25. **a.** 720 kPa
  b. about 7 times higher
26. 1.1 atm
27. 136 K
28. −150°C
29. 273°C
30. 273°C

Answers to Caption Questions

Figure 11.1 Atoms of hydrogen and atoms of chlorine have different electronegativities, causing a permanent dipole effect.

Figure 11.5 Increasing temperature causes the average kinetic energy of molecules to increase. They strike the surface of a container more often and harder, causing the volume of a flexible container to increase.

Figure 11.7 As atmospheric pressure changes, the level of mercury in the tube changes. Greater pressure causes the mercury column to rise and lower atmospheric pressure causes the column to lower.

Figure 11.14 If the Celsius temperature is doubled, the Kelvin temperature is not doubled. Instead it goes up by the amount of the difference between the two Celsius temperatures.

Figure 11.15 The *x*-intercept for the Celsius graph is −273 °C, and the *y*-intercept varies depending on initial pressure and volume. The *x*-and *y*-intercept for the Kelvin graph are both 0.

Figure 11.17 A relief valve is a safety device to prevent the cylinder from exploding. At a critical pressure the relief valve opens and permits the escape of gas, so that the pressure inside the tank does not increase to dangerous levels.

Answers to Section 11.2 Review Questions

5. **a.** 1.25 atm
  b. 1.5 bar
  c. 105 kPa
  d. 1.25 atm
11. 3.0 L
12. 417 kPa
13. 1000

Answers to Section 11.3 Review Questions

3. **a.** 311 K
  b. 395.6 K
  c. 233 K
  d. 1.85°C
  e. −99.55°C
  f. 6.00×10^2°C
8. 0.7 L
9. doubling of the temperature, to 60°C
11. 75.4 kPa
12. 6500 kPa
13. 17.4 psi

Answers to Chapter 11 Review Questions

1. c
2. d
3. d
4. b
5. c
6. b
7. e
8. e

13. a. 551 kPa
  b. 0.2 atm to 0.1 atm
  c. 80 kPa
  d. 1520 mmHg
  e. 1.2 atm
  f. 101 kPa

15. 214 kPa

16. 0.16 L

18. a. 298 K
  b. 263.2 K
  c. 3.85°C
  d. −108°C

19. −25°C

20. b. 303 kPa

22. 923°C

24. 32 mL; −93°C

27. a. 59°C
  b. 175 cm^3

29. c. 1.82×10^3 cm^3; the volume changes by a factor of 1.52

Answers to Chapter 11 Self-Assessment Questions

1. c
2. a
3. e
4. a
5. a
6. a
7. b
8. a
9. b
10. c

16. a. 79.9 psi
  b. 310 mmHg
  c. 53 kPa
  d. 677 mmHg; 90.2 kPa

17. a. $x = 1.7 \times 10^3$ mL

18. 100 kPa

21. 2.1×10^2 mL

22. 3.1 L

23. 40 000 kPa

24. 11.7 psi

Chapter 12

Answers to Learning Check Questions

1. Boyle's and Charles's laws

2. Volumes of gaseous reactants and products are always in whole-number ratios, when measured at the same temperature and pressure.

3. 586 mL

4. The ratio of the volumes is the same as the ratios of the coefficients of the balanced equation.

5. b. 62 mL

6. 15 L

7. combined gas law and Avogadro's law

8. For using the universal gas constant value of 8.31 kPa•L/mol•K, the units to use are: pressure in KPa, temperature in K, and volume in L.

9. 275 kPa

10. helium: 0.24 atm; argon: 0.50 atm; neon: 1.3 atm

11. Humidity represents water vapour in the air, which contributes its own partial pressure to the total pressure.

12. 81.1 kPa

13. In the homosphere layer the gases are evenly mixed and behave like an ideal gas. In the heterosphere the gases are layered and more limited and include charged ionic gases.

14. The main gases present are nitrogen, oxygen, argon and carbon dioxide.

15. Criteria air contaminants are those pollutant gases that are considered to have the greatest impact on air quality and human health.

16. Convection occurs in the homosphere due to heating of the land by radiant energy. As gases are heated at ground level, they become less dense and will rise upwards while the colder, denser gases sink downwards.

17. AQHI stands for Air Quality Health Index. It is a scale used to indicate air quality and associated health risks.

18. Air quality in Ontario would reach levels over 7 when temperatures are high and air is not mixing. This would occur mostly in summer months and during periods of heavy traffic and industrial emissions.

Answers to Practice Problems

1. 620°C

2. 13.6 ml

3. 0.214 mL

4. 22°C

5. 104 kPa

6. 1.2 atm

7. 21 mL

8. 57°C

9. 3.7×10^3 cm^3

10. 6×10^3 m^3

11. 1.6 g

12. 25 L

13. 17.7 L

14. 1.69 L

15. 141 L

16. 8.58 g; 3.03×10^{23} molecules

17. 26 L/mol

18. 25.9 L/mol

19. a. 1.46×10^{-3} mol
  b. 1.46×10^{-3} mol
  c. 24.5 L/mol

20. 22.4 L/mol

21. 140 L

22. 11 L

23. 166°C

24. 118 kPa

25. 160 g

26. 60 g/mol

27. 83.8 g/mol; element is most likely krypton

28. 1.77 g/L

29. a. C_2H_5
  b. 58.1 g/mol
  c. C_4H_{10}

30. C_8H_{18}

31. 36 mL

32. 16 L

33. 40 L

34. 5.40×10^2 L

35. 15.6 L

36. 0.787 L

37. 2.73×10^{-3}g

38. 0.32 g

39. 0.17 g

40. 0.16 L

Answers to Caption Questions

Figure 12.5 The pressure would be 2.0 atm instead of 1.5 atm.

Figure 12.6 If any air were already in the cylinder before the gas was collected, the measured volume of the collected gas would be incorrect.

Figure 12.9 Gay-Lussac's law; could be placed connecting to the combined gas law

Answers to Section 12.1 Review Questions

4. 84°C

5. 74.8 kPa

6. 4×10^3 hours

7. 29 L

8. 30 L

9. 7.0×10^2 mL

11. 0.14 mol

12. a. 0.446 moles
  b. 17.8 g
  c. 2.69×10^{23} atoms
  d. 22.4 L/mol

14. 42.8 L/mol

Answers to Section 12.2 Review Questions

2. a. $V = \frac{nRT}{P}$
  b. $P = \frac{nRT}{V}$
  c. $T = \frac{PV}{nR}$
  d. $n = \frac{PV}{RT}$

3. 8.6×10^3 L

4. 8.3 kPa

5. 24.2 L/mol

6. 38.1 g/mol, F_2

7. 1.2 g/L

8. C_2H_2

9. 97.5 kPa

11. 109 g

12. 1.43 L

13. 3.2 L

Answers to Chapter 12 Review Questions

1. a
2. c
3. b
4. b
5. e
6. c
7. a
8. d

11. 127 L

12. 4900 kPa

13. 900 kPa

14. 2.75 g

15. 1.0×10^2 kPa

16. b. 1.2 mol
  c. 32 g

17. a. 140 g/mol
  b. 120 g/mol

18. 20.2 g/mol; neon

19. $C_4H_{12}O_2$

20. 1.4 g/L

21. 97.5 kPa

22. a. $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$
  b. 2.5×10^3 L
  c. 2.0×10^3 L
  d. 3.0×10^3 L
  e. 5.0×10^3 L

23. a. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
  b. 1 L
  c. 3 L
  d. 4 L

24. a. $NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(g)$
  b. 0.30 g

25. 69 L

26. 0.25 g Mg; 6.7 mL HCl

34. a. 2.2×10^6 g
  b. 7.1×10^6 g

Answers to Chapter 12 Self-Assessment Questions

1. c
2. a
3. e
4. d
5. d
6. e
7. c
8. b
9. d
10. b

12. 20 m^3

14. 24.4 L/mol

15. a. 1.1×10^6 g
  b. 8.9×10^6 L

17. molar mass 2.0 g/mol; H_2

18. 30.5 kPa

20. $2N_2(g) + O_2(g) \rightarrow 2N_2O(g)$; volume of $N_2(g)$: 34 L; volume of $O_2(g)$ 17 L

21. 6.0×10^3 g

Answers to Unit 5 Review Questions

1. b
2. e
3. e
4. b
5. e
6. c
7. e
8. b
9. b
10. d

18.

Pressure Unit Conversions

Pressure (kPa)	Pressure (mmHg) or Torr	Pressure (atm)	Pressure (PSI) lbs/inch
10^2	765	1.01	14.8
3.0×10^2	2300	3.0	44
3.00×10^2	2.07×10^3	3.00	40.0

20. 617 cm^3

21. a. 12 psi

22. 22 atm

23. −18°C

24. a. 2.8 g/L
  b. 0.50 g/L
  c. 0.635 g/L

25. 1.2 L/mol

26. 1.8×10^4 kg

27. 1.3×10^4 kPa

28. a. 0.011 mol
  b. 0.50 g
  c. 6.8×10^{21}

29. Molar mass is 38 g/mol and identity is F_2

30. a. 98 kPa
  b. 4.3×10^{-3} mol
  c. 4.3×10^{-3} mol
  d. 88 g/mol
  e. *X* is magnesium. The carbonate formula is $MgCO_3$.

31. a. empirical formula: C_1H_1
  b. molar mass: 78 g/mol
  c. molecular formula: C_6H_6

32. a. $3CO(g) + 7H_2(g) \longrightarrow C_3H_8(g) + 3H_2O(g)$
  b. 1.050×10^3 L
  c. 466.7 L
  d. 38.6 L

33. mass of $CaCO_3(s)$ in tablet: 0. 40 g; 8.0×10^1% (m/m)

34. 180 L

35. a. Ammonium nitrate produces 310 L of nitrous oxide gas, while ammonia produces 730 L. The volume of nitrous oxide produced from ammonia is over twice that produced by the same mass of ammonium nitrate.
  b. 2500 L $N_2O(g)$, 7500 L $H_2O(s)$
  c. 4.8×10^5 g

36. a. 2.7×10^4 L
  b. 1.2×10^4 L
  c. 3.6×10^4 L
  d. 7.5×10^4 L

37. a. lithium hydride reaction: $LiH(s) + H_2O(\ell) \longrightarrow LiOH(aq) + H_2(g)$; magnesium hydride reaction: $MgH_2(s) + 2H_2O(\ell) \longrightarrow Mg(OH)_2(aq) + 2H_2(g)$
  b. 33 g LiH(s), 220 g $MgH_2(s)$

50. b. lung volume at a depth of 500.0 m: 1.1×10^2 L; lung volume at a depth of 1.0×10^3 L: 51 L

Answers to Unit 5 Self-Assessment Questions

1. a
2. a
3. c
4. c
5. c
6. b
7. d
8. e
9. e
10. d

13. a. height of liquid gallium: 1.68×10^3 mmHg

14. 27 L

16. 21°C

18. 7.0×10^3 kPa

19. b. 1.2 mol
  c. 35 g
  d. 7.5×10^{23} molecules

20. 2700 cm_3

22. a. 97.7 kPa
  b. 0.010 mol $CO_2(g)$
  c. 0.010 mol $X_2CO_3(s)$
  d. molar mass of carbonate: 1.4×10^2 g/mol (137.8 g/mol — rounded to appropriate number of significant digits)
  e. The element represented by X is potassium.