PASS CHEMISTRY BOOK 1500



Lindsay Blackstock & Sharon Brewer
Thompson Rivers University

LibreTexts™

Thompson Rivers University

PASS Chemistry Book 1500

Lindsay Blackstock & Sharon Brewer

This text is disseminated via the Open Education Resource (OER) LibreTexts Project (https://LibreTexts.org) and like the thousands of other texts available within this powerful platform, it is freely available for reading, printing, and "consuming."

The LibreTexts mission is to bring together students, faculty, and scholars in a collaborative effort to provide and accessible, and comprehensive platform that empowers our community to develop, curate, adapt, and adopt openly licensed resources and technologies; through these efforts we can reduce the financial burden born from traditional educational resource costs, ensuring education is more accessible for students and communities worldwide.

Most, but not all, pages in the library have licenses that may allow individuals to make changes, save, and print this book. Carefully consult the applicable license(s) before pursuing such effects. Instructors can adopt existing LibreTexts texts or Remix them to quickly build course-specific resources to meet the needs of their students. Unlike traditional textbooks, LibreTexts' web based origins allow powerful integration of advanced features and new technologies to support learning.



LibreTexts is the adaptable, user-friendly non-profit open education resource platform that educators trust for creating, customizing, and sharing accessible, interactive textbooks, adaptive homework, and ancillary materials. We collaborate with individuals and organizations to champion open education initiatives, support institutional publishing programs, drive curriculum development projects, and more.

The LibreTexts libraries are Powered by NICE CXone Expert and was supported by the Department of Education Open Textbook Pilot Project, the California Education Learning Lab, the UC Davis Office of the Provost, the UC Davis Library, the California State University Affordable Learning Solutions Program, and Merlot. This material is based upon work supported by the National Science Foundation under Grant No. 1246120, 1525057, and 1413739.

Any opinions, findings, and conclusions or recommendations expressed in this material are those of the author(s) and do not necessarily reflect the views of the National Science Foundation nor the US Department of Education.

Have questions or comments? For information about adoptions or adaptions contact info@LibreTexts.org or visit our main website at https://LibreTexts.org.

This text was compiled on 06/08/2025

TABLE OF CONTENTS

Licensing

1: Introduction

- 1.1: About PASS and the Creators
- 1.2: How to use PASS Chem

2: Quantum Theory and Electronic Structure

- 2.1: Question 2.E.06 PASS - Energy, wavelength, frequency, and colour of emitted lithium photons
- 2.2: Question 2.E.26 PASS - Bohr Model, quantized energy change
- 2.3: Question 2.E.40 PASS - quantum numbers, orbital characteristics
- 2.4: Question 2.E.58 PASS - electron configuration, orbital diagrams, Aufbau's Principle Hund's Rule

3: Periodic Relationships Among the Elements

- 3.1: Question 3.E.05 PASS - ranking first ionization energy
- 3.2: Question 3.E.09 PASS - ranking atomic radius
- 3.3: Question 3.E.18 PASS - describe ionic radii of isoelectronic series

4: Chemical Bonding I - Basic Concepts

- 4.1: Question 4.E.20 PASS - which is the most polar bond?
- 4.2: Question 4.E.60 PASS - draw Lewis Structure and calculate formal charge
- 4.3: Question 4.E.65 PASS - determining enthalpy change from bond energies
- PASS Lewis Structure H5P Activity Interactive Video (4.E.60)

5: Chemical Bonding II - Molecular Geometry and Hybridization of Atomic Orbitals

- 5.1: Question 5.E.09 PASS - determine electron pair geometry and molecular geometry
- 5.2: Question 5.E.17 PASS - which central atom makes a polar molecule?
- 5.3: Question 5.E.59 PASS - identify central atom hybridization
- 5.4: Question 5.E.78 PASS - using molecular orbital theory and bond order to determine ion charge

6: Intermolecular Forces and Liquids and Solids

- 6.1: Question 6.E.07 PASS - define and explain the types of intermolecular forces
- 6.2: Question 6.E.11 PASS - arrange in order of increasing boiling point
- 6.3: Question 6.E.15 PASS - predict melting point

7: Organic Chemistry I - Bonding and Structure

- 7.1: Question 7.1.7 PASS - label hybridized atomic orbitals, identify sigma and pi bonds
- 7.2: Question 7.2 PASS - interpret chemical structure line drawing
- 7.3: Question 7.4.6 PASS - identifying constitutional isomers
- 7.4: Question 7.5.2.1 PASS - naming cycloalkane structures

8: Organic Chemistry II - Stereochemistry

- 8.1: Question 8.5.2.3 PASS - alkenes, label E or Z
- 8.2. Question 8.4.E.1 PASS - stereocenter configuration R or S
- 8.4 Question 8.4.E.2 PASS - enantiomers, diastereomers, identical

9: Organic Chemistry III - Conformational Analysis

- 9.1: Question 9.E.1 PASS - butane Newman projection, bond rotation energy diagram
- 9.2: Question 9.E.2 PASS - cycloalkane chair, axial vs. equatorial, most stable

Index

Glossary

Detailed Licensing

Detailed Licensing

Licensing

*A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.*

CHAPTER OVERVIEW

1: Introduction

1.1: About PASS and the Creators
1.2: How to use PASS Chem

1: Introduction is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.1: About PASS and the Creators

The PASS (Platform Adapted Strategic Solutions) Framework was developed to provide a literature-based strategy for OER practice problem solutions for general chemistry practice questions. The PASS solutions presented in these collections contain templated solutions developed for OER chemistry problems across first year chemistry topics at Thompson Rivers University (TRU) that were adapted for the LibreText platform.



The content creators pictured above are Assistant Teaching Professor Dr. Lindsay Blackstock, Student Research Partner Ashlynn Jensen, and Associate Professor Dr. Sharon Brewer. The content development was made possible by funding from a TRU Instructional Innovation Grant (TIIG) awarded to Dr. Blackstock & Dr. Brewer in 2023. The PASS framework development was funded by a TRU Open Educational Resources Grant awarded to Dr. Blackstock, Dr. Brewer and Dr. Verena Roberts in 2021.

1.1: About PASS and the Creators is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.2: How to use PASS Chem

This book follows the Platform Adaptable Strategic Solution (PASS) template designed by Drs. Lindsay Blackstock, Sharon Brewer, and Verena Roberts to support students at various levels of understanding and mimic the experience that a student may have during an office hour. You can find the first examples using this template on the PASS Chem website: https://chempass.opened.ca/.

For each question, students may attempt and check their answer with **Show/Hide Answer.** If you were correct, **AMAZING!** You can move on to the next question.

Strategy Map: If you need assistance in *what general steps to take* (i.e., spoiler-free!) then move on to the **strategy map**. The strategy map will provide you with a procedure (instructions) of what to do to reach the answer.

Solution: If you would like to see each step (i.e., spoiler alert!), move on to the **solution**. This section shows a simple step-by-step solution to reach the answer but without any commentary.

Guided Solution: If you need more help and want to know *why* you are performing each step in the solution, move on to the **guided solution**.

The guided solution explains the chemistry theory or rationale behind the process. The guided solution's purpose is to replicate what a student may experience during an office hour or a lecture, working through a problem supported by the 'teacher talk.' Here, you will also find supplemental information including hints embedded throughout, such as:

- Links to read more from the LibreTexts reference textbooks
- Additional context or an alternative perspective
- Reminders about various unit conversions or significant figures
- Common mistakes to avoid

Complete Solution: Below the guided solution is the **complete solution** that combines the step-by-step solution with the key context points from the guided solution.

Check Your Work: In chemistry, you can often avoid mistakes by thinking critically about the theory related to the problem and predicting what the answer *should or shouldn't be*. Be sure to **check your work** (see the purple box below). This section provides the rationale of why your answer 'makes sense' with respect to the chemistry theory. Use this strategy to avoid making mistakes in similar problems.

Our suggested approach:

1. Attempt it yourself. Check to see if you need additional assistance to get the correct answer.
2. After checking your answer, decide how much additional support you need to understand how to solve the problem yourself. Navigate to one of the three boxes below:
    1. **Strategy Map**
    2. **Solution**
    3. **Guided Solution.**

    The boxes are listed in order of how much assistance is provided from least to most.

Below the answer is the link to the online resource to find supplementary information about this topic.

We hope this helps you build your confidence in working through these practice chemistry problems!

1.2: How to use PASS Chem is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

2: Quantum Theory and Electronic Structure

2.1: Question 2.E.06 PASS - Energy, wavelength, frequency, and colour of emitted lithium photons

2.2: Question 2.E.26 PASS - Bohr Model, quantized energy change

2.3: Question 2.E.40 PASS - quantum numbers, orbital characteristics

2.4: Question 2.E.58 PASS - electron configuration, orbital diagrams, Aufbau's Principle Hund's Rule

---

2: Quantum Theory and Electronic Structure is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.1: Question 2.E.06 PASS - Energy, wavelength, frequency, and colour of emitted lithium photons

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 2.1.2.E.06

Heated lithium atoms emit photons of light with an energy of 2.961 ×10^{-19} J. Calculate the frequency and wavelength of one of these photons. What is the total energy in 1 mole of these photons? What is the color of the emitted light?

Answer

$$\begin{aligned}\text{Frequency (v)} &= 4.469 \times 10^{14}\ \text{s}^{-1}\\ \text{Wavelength } (\lambda) &= 6.713 \times 10^{-7}\ \text{m}\\ \text{Energy in 1 mole } &= 1.783 \times 10^{5}\ \text{J}\\ \text{Colour of light} &= \text{ Red}\end{aligned}$$

> See LibreText Section 2.1 Electromagnetic Energy
>
> See LibreText Example 2.1.2 Calculating the Energy of Radiation

Strategy Map

Step	Hint
1. Identify the information provided in the question.	Recall: LibreText Example 2.1.2 Calculating the Energy of Radiation
2. Choose a formula that connects the information that is provided and the frequency and wavelength.	You will need to use two formulas
3. Using provided data to calculate the amount of energy in one mole.	How many **"items"** are in one mole? How can you convert a **photon** to a **mole**?
4. Compare your wavelength to the electromagnetic spectrum, what colour would your calculated wavelength fall under?	Try converting your wavelength into nanometers.

Solution

Frequency of emitted photon

$$E = hv$$
$$2.691 \times 10^{-19} J = \left(6.626 \times 10^{-34} Js\right) v$$
$$v = \frac{6.691 \times 10^{-19} J}{6.626 \times 10^{-34} Js}$$
$$v = \mathbf{4.469 \times 10^{14}} \mathbf{s}^{-1}$$

Wavelength of emitted photon

$$c = \lambda v$$
$$3.00 \times 10^{8}\ \text{m/s} = \left(4.469 \times 10^{14}\ \text{s}^{-1}\right) \lambda$$
$$\lambda = \frac{3.00 \times 10^{8}\ \text{m/s}}{4.469 \times 10^{14}\ \text{s}^{-1}}$$
$$\lambda = \mathbf{6.713 \times 10^{-7}} \text{m}$$

Energy in 1 mole of emitted photons

$$\frac{2.961 \times 10^{-19}\ \text{J}}{1\ \text{photon}} \times \frac{6.022 \times 10^{23}\ \text{photons}}{1\ \text{mole}}$$
$$= \mathbf{1.783 \times 10^{5}\,J/mol}$$

Colour of emitted light

$$6.713 \times 10^{-7}\ \text{m} \times \frac{1 \times 10^{9}\ \text{nm}}{1\ \text{m}} = 671\ \text{nm}$$
$$= \text{Red Light}$$

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a multiple part calculation question where you use related equations to solve for frequency and wavelength. Next to determine secondary information use constant values and to calculate the energy in one mole of the given photons and conversion factors to predict the colour of light emitted.	Refer to: LibreText Example 2.1.2 Calculating the Energy of Radiation
Highlight important information provided in the question: Heated lithium atoms emit photons of light with an energy of 2.961×10^{-19} J. Calculate the frequency and wavelength of one of these photons. What is the total energy in 1 mole of these photons? What is the color of the emitted light?"	When reading through a question to know what to find look for key words such as "Calculate" and phrases such as "What is". - "Calculate the frequency and wavelength of one of these photons." - "What is the total energy in 1 mole of these photons?" - "What is the color of the emitted light?"
An example of the same question with key information emphasized: Heated lithium atoms emit photons of light with an **energy** of **2.961×10^{-19} J**. • *Calculate* the **frequency** and **wavelength** of one of these photons. • *What is* the **total energy** in **1 mole** of these photons? • *What is* **the color** of the emitted light?	A summary of the key information provided: What we know: We know that the energy (E) of an emitted photon is equal to 2.961×10^{-19} **J**. What we are asked: We are asked to find the **frequency** (v), **wavelength** (λ), the **energy in one mole,** and **the colour** of the emitted light.
What else do you know about this topic? Avogadro's Number defines the number of 'items' in a mole Visual Light spectrum: • Violet: 400-420 nm • Indigo: 420-440 nm • Blue: 440-490 nm • Green: 490-570 nm • Yellow: 570-585 nm • Orange: 585-620 nm • Red: 620-780 nm	**Recall that** you know how many particles are in one mole. (Avogadro's Number = 2.066×10^{23}) Photons are also known as particles. We can use this value in a **conversion** calculation to convert our photons into one mole of photons. **Recall that** each colour on the light spectrum is seen as a different wavelength, typically seen in nanometers. To find the colour of light convert your wavelength into nanometers.

Guided Solution	Hint
Complete Solution: **Step 1** Using Planck's equation, we can rearrange and solve for the wave's frequency (***v***). We are given the energy and we know Planck's constant (***h*** = 6.626x10^{-34} Js) $$E = hv$$ $$2.691 \times 10^{-19}\,J = \left(6.626 \times 10^{-34}\,Js\right) v$$ $$v = \frac{6.691 \times 10^{-19}\,J}{6.626 \times 10^{-34}\,Js}$$ $$v = \mathbf{4.469 \times 10^{14}\,s^{-1}}$$ **Step 2** We can now rearrange and solve for our wavelength (**λ**) using the frequency we just found and the speed of light (c = 3.00x10^{8} m/s) in the equation c = λv. $$c = \lambda v$$ $$3.00 \times 10^{8}\ \mathrm{m/s} = \left(4.469 \times 10^{14}\ \mathrm{s^{-1}}\right) \lambda$$ $$\lambda = \frac{3.00 \times 10^{8}\ \mathrm{m/s}}{4.469 \times 10^{14}\ \mathrm{s^{-1}}}$$ $$\lambda = \mathbf{6.713 \times 10^{-7}\,m}$$ **Step 3** To find the energy in one mole of these photons we need to use Avogadro's number as a conversion factor. There are 2.066x10^{23} photons in one mole. $$\frac{2.961 \times 10^{-19}\ \mathrm{J}}{1\ \mathrm{photon}} \times \frac{6.022 \times 10^{23}\ \mathrm{photons}}{1\ \mathrm{mole}}$$ $$= \mathbf{1.783 \times 10^{5}\,J/mol}$$ **Step 4** To find the colour of light emitted, first convert from meters to nanometers (this makes it easier to compare the calculated value to known colour ranges in the visible light spectrum). $$6.713 \times 10^{-7}\ \mathrm{m} \times \frac{1 \times 10^{9}\ \mathrm{nm}}{1\ \mathrm{m}} = 671\ \mathrm{nm}$$ $$= \text{Red Light}$$ The calculated value falls in the range of red light which is between 620 and 780 nm.	Re: Step 3 - To set up a conversion properly, be sure that the numerator has the unit you want, and the denominator has the unit that you must cancel out. Re: Step 4 - Recall that each colour on the light spectrum has a different wavelength, typically described in nanometers . To find the colour of light convert your wavelength into nanometers. There are 10^{9} nanometers in 1 meter; alternately 1 nanometer is 10^{-9} meters. Either conversion factor will work, just be sure that the units cancel out appropriately. Violet: 400-420 nm; Indigo: 420-440 nm; Blue: 440-490 nm; Green: 490-570 nm; Yellow: 570-585 nm; Orange: 585-620 nm; Red: 620-780 nm
Check your work! **Why does this answer make chemical sense?** Wavelength and frequency are inversely proportional to one another, as frequency increases, wavelength decreases. This is demonstrated in our answer as the wavelength is very small and the frequency is very large. The colour of light is dependent on the wavelength of the traveling light. Different wavelengths will be absorbed and reflected differently producing different colours in the visible light spectrum. In this case the wavelength was on the higher end of that spectrum and thus produced red light.	

(question source from LibreText page titled 6.E:Electronic Structure and Periodic Properties (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/06%3A_Electronic_Structure_and_Periodic_Properties_of_Elements/6.E%3A_Electronic_Structure_and_Periodic_Properties_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry-2e/pages/6-exercises , Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

2.1: Question 2.E.06 PASS - Energy, wavelength, frequency, and colour of emitted lithium photons is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.2: Question 2.E.26 PASS - Bohr Model, quantized energy change

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 2.2.2.E.26

Using the Bohr model,

a. determine the change in energy (in joules) when the electron in a Li^{2+} ion moves from the n = 2 to the n = 1 orbit.

b. What is the energy (in joules) of the photon produced?

Answer

$$a.\ \Delta E = -1.47 \times 10^{-17}\ J \tag{2.2.1}$$

$$b.\ E_{photon} = 1.47 \times 10^{-17}\ J \tag{2.2.2}$$

> See the Bohr Model (see LibreText section 2.2)

Strategy Map

Step	Hint
1. Identify the information (data and/or variables) given in the question.	> Recall the Bohr Model (see LibreText section 2.2) n_i=2, n_f=1 (**i**nitial and **f**inal energy levels)
2. Identify what the question is asking for; what variable would represent it in a formula.	Looking for ΔE (total energy change)
3. Choose a formula that allows you to plug in your given information. Fill in any known constant values.	$$\Delta E = -2.18 \times 10^{-18} J \left(\frac{Z^2}{n_f^2} - \frac{Z^2}{n_i^2} \right) \tag{2.2.3}$$ to check if using the correct formula; does this formula help you? Are you closer to **finding** the desired variable? n_i=2 n_f=1 ΔE=? Rydberg's Constant (R_H)=2.18x10^{-17} Z=Atomic Number=3 **Note that the atomic number is found on the Periodic Table, most questions consider 'hydrogen's electron'; because hydrogen has an atomic number of Z=1, the formula is often simplified to: $$\Delta E = -R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \tag{2.2.4}$$

Step	Hint
4. Calculate the final answer, do not forget about sign, units and significant figures.	Be sure to follow typical math rules such as order of operations > Recall the Order of Operations (see LibreText section 1.4) **Common mistakes** are made when filling in the formula. When using nonhydrogen atoms, students often forget to include the atomic number (Z) in the numerator of both fractions. In this example, the atomic number is for Li^{2+} is 3.

Solution

$$a.\ \Delta E = -R_H \left(\frac{Z^2}{n_f^2} - \frac{Z^2}{n_i^2} \right) \tag{2.2.5}$$

$$\Delta E = -2.18 \times 10^{-18}\ J \left(\frac{Z^2}{n_f^2} - \frac{Z^2}{n_i^2} \right)$$

$$\Delta E = -2.18 \times 10^{-18}\ J \left(\frac{3^2}{1^2} - \frac{3^2}{2^2} \right)$$

$$\Delta E = -2.18 \times 10^{-18}\ J \left(\frac{9}{1} - \frac{9}{4} \right)$$

$$\Delta E = -2.18 \times 10^{-18}\ J \left(\frac{27}{4} \right)$$

$$\Delta E = -2.18 \times 10^{-18}\ J (6.25)$$

$$\Delta E = -1.47_{15} \times 10^{-17}\ J$$

$$\Delta E = -1.47 \times 10^{-17}\ J$$

$$b.\ E_{photon} = 1.47 \times 10^{-17}\ J \tag{2.2.6}$$

Guided Solution

Guided Solution	Hint
This is a calculation question where you use the Rydberg Equation to calculate the energy of the transition of the electron.	Recall the Bohr Model (see LibreText Example 2.2.1)
Using the Bohr model, a. determine the **change in energy** (in joules) when the electron in a Li^{2+} ion moves **from** the n = 2 **to** the n = 1 orbit. b. What is the **energy** (in joules) of the photon produced?	change in energy = ΔE "from the" = initial energy level = n_i = 2 "to the" = final energy level = n_f =1 energy of the photon = E_{photon} = |ΔE|
Summarize what the question is asking	we are looking for the change in energy of the Li^{2+} electron when it transitions from n=2 to n=1. The energy change is exactly the same (AKA quantized) amount as what is emitted (i.e., as a photon).

	Guided Solution	Hint
tion	Predict what you expect the answer to be based on your understanding of the related chemistry concepts	Since the Li^{2+} ion is going from a higher energy orbital (n=2) to a lower energy orbital (n=1), you can expect the answer to be a negative value. • When energy is emitted is it released or lost from the system; ΔE is a negative value. • When energy is absorbed, it is added to the system; ΔE is a positive value
tion	**Complete Solution:** **PART a.** **Step 1** Recall the Rydberg equation: $$\Delta E = -R_H \left(\frac{Z^2}{n_f^2} - \frac{Z^2}{n_i^2} \right) \quad (2.2.7)$$ **Step 2** Input your data provided by the question: • ΔE = unknown • Z = atomic number of atom (Li^{2+}) = 3 • n_f = Final energy state = 1 • n_i = Initial energy state = 2 • R_H = Rydberg's Constant (2.18x10^{-17} J) $$\Delta E = -2.18 \times 10^{-18}\ \text{J} \left(\frac{3^2}{1^2} - \frac{3^2}{2^2} \right) \quad (2.2.8)$$ **Step 3** Start by solving inside the brackets: $$\Delta E = -2.18 \times 10^{-18}\ \text{J} \left(\frac{9}{1} - \frac{9}{4} \right) \quad (2.2.9)$$ $$\Delta E = -2.18 \times 10^{-18}\ \text{J} \left(\frac{27}{4} \right)$$ $$\Delta E = -2.18 \times 10^{-18}\ \text{J}(6.25)$$ **Step 4** Multiply the constant (R_H) and the value inside the brackets: $$\Delta E = -1.47_{15} \times 10^{-17}\ \text{J} \quad (2.2.10)$$ $$\Delta E = -1.47 \times 10^{-17}\ \text{J}$$ **Part b.** **Step 5** Recall the relationship between the energy change for the transition of an electron to the energy of the photon that is absorbed or emitted during that change energy of the photon = $E_{photon} = \lvert\Delta E_{electron}\rvert$ $\Delta E_{electron}$ = -1.47x10^{-17} J therefore E_{photon} = 1.47x10^{-17} J	Re: Step 1 - the Rydberg equation allows us to calculate the **overall change in energy** (ΔE) of an atom, given its **initial** (n_i) and **final** (n_f) **energy values**. Re: Step 2 - **Common mistakes** are made here, when using nonhydrogen atoms, students often forget to include the atomic number (Z) as the numerator. Re: Step 3 - Square the numerator and denominator of both fractions. Combine the fractions by subtracting them. An optional step, to convert your fraction to a decimal before the multiplication step. Re: Step 4 - notice that there is a negative sign in front of the constant value. Re: Step 5 - the energy of a photon must exactly match the energy change resulting from the transition of the electron from one energy level (orbit) to another (orbit) A photon is a quantized packet of energy When an electron transitions from a low to a high energy level, the electron must absorb energy ($\Delta E_{electron}$ = positive) from a photon with the exact amount of energy required for that transition When an electron transitions from a high to a low energy level, the electron must release energy ($\Delta E_{electron}$ = negative). The energy lost from the electron is emitted as a photon with the exact amount of energy as was released during that transition.

Guided Solution	Hint
Check your work! Why does this answer make chemical sense?	The calculated answer for this problem was a negative value. This number makes does make chemical sense as our final energy state is smaller than our initial energy state, ($1/n_{final} - 1/n_{initial}$) a small number subtracted from a larger number results in a positive value inside the brackets which is then multiplied by the negative Rydberg's constant. Because the change in energy was negative this means the energy was emitted. Recall that when atoms go from a higher energy state to a lower energy state, they emit a photon of electromagnetic energy, that amount of energy is dependent on the difference of energy in the two states. Emitted energy relates to a negative calculation, whereas absorbed energy relates to a positive calculation.

(question source adapted question 6.2.11 from 6.E: Electronic Structure and Periodic Properties (Exercises) - Chemistry LibreTexts, shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source question 27 in Ch. 6 Exercises - Chemistry 2e | OpenStax, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

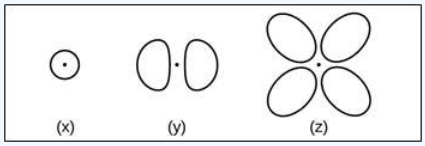
2.2: Question 2.E.26 PASS - Bohr Model, quantized energy change is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.3: Question 2.E.40 PASS - quantum numbers, orbital characteristics

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 2.3.2.*E*.40

Consider the orbitals shown here in outline…



1. What is the maximum number of electrons contained in an orbital of type (x)? Of type (y)? Of type (z)?
2. How many orbitals of type (x) are found in a shell with n = 2? How many of type (y)? How many of type (z)?
3. Write a set of quantum numbers for an electron in an orbital of type (x) in a shell with n = 4. Of an orbital of type (y) in a shell with n = 2. Of an orbital of type (z) in a shell with n = 3.
4. What is the smallest possible n value for an orbital of type (x)? Of type (y)? Of type (z)?
5. What are the possible l and ml values for an orbital of type (x)? Of type (y)? Of type (z)?

Answer

1. type (x) = 2 electrons; type (y) = 2 electrons; type (z) = 2 electrons
2. type (x) = 1 orbital; type (y) = 3 orbitals; type (z) = 0 orbitals
3. type (x) = [4, 0, 0, ½]; type (y) = [2, 1, 0, -½]; type (z) = [3, 2, -1, ½]
4. type (x): n = 1; type (y): n = 2; type (z): n = 3
5. type (x): l = 0, ml = 0; type (y): l = 1, ml = -1, 0, 1; type (z): l = 2; ml = -2, -1, 0, 1, 2

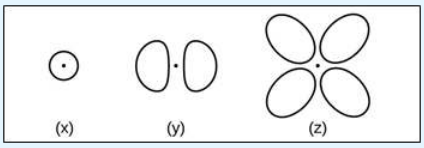
> See LibreText 2.4 Quantum Mechanics and the Atom

Strategy Map

Step	Hint
This problem covers concepts from several topics in section 2: Quantum Theory and Electronic Structure	Refer to: • For a visual representation of the orbital types and their suborbital positions see figure 2.5.7. Libre Text 2.6, Electronic Structure of Atoms (Electron Configurations) • For a visual representation of the hierarchy of subshell filling/energy levels see figure 2.6.3.
1. Identify what orbital types are represented in the figure.	In CHEM1500 four types of orbitals are introduced: s, p, d, and f. Recall this mnemonic the letter can help describe the shape • s (sphere - 1 lobe) • p (propeller - 2 lobes) • d (daisy - 4 lobes; or donut - two lobes and a ring shape) • f (funky - 7 lobes)

Step	Hint
2. Consider each requirement of this multi-part problem and work through the answer step by step to not miss anything	Be familiar with the fundamental concepts of atomic orbital theory and the key characteristics of each orbital type. Consider Hund's Rule, Aufbau's Principal, and the Pauli Exclusion Principal to guide your thinking.

Guided Solution

Guided Solution	Hint
This problem requires comprehension of the following concepts: 'quantum mechanics and the atom', 'shapes of atomic orbitals', and 'electronic structure of atoms' Consider the orbitals shown here in outline…  	Refer to: • Libre Text 2.4, Quantum Mechanics and The Atom • Libre Text 2.5, The Shape of Atomic Orbitals • Libre Text 2.6, Electronic Structure of Atoms (Electron Configurations)

Recall the quantum numbers:
• what are they and what do they represent?
  ◦ what are the guidelines/rules/principals that inform each quantum number?
	Quantum numbers: [n, l, ml, ms]
	• n (principal quantum number; represents shell energy level)
	  ◦ must be a whole number integer (i.e., n = 1, 2, 3, ... n)
	  ◦ bigger number = greater energy = greater distance from nucleus = larger size
	• l (azimuthal / orbital angular momentum quantum number, represents orbital shape)
	  ◦ must be zero or a positive whole number integer, restricted by n (i.e., l = 0, 1, 2, ... n-1; l≠n)
	  ▪ when l=0, type **s** (sphere shaped: 1 lobe)
	  ▪ when l=1, type **p** (propeller shaped: 2 lobes)
	  ▪ when l=2, type **d** (daisy shaped: 4 lobes; or donut shaped: two lobes and a ring shape)
	  ▪ when l=3, type **f** (funky shaped: 7 lobes)
	• ml (magnetic quantum number, represents the orientation of the orbital within a subshell)
	  ◦ must be a positive or negative whole number integer, or zero; restricted by l (i.e., ml = -l, ... -2, -1, 0, 1, 2, ... +l)
	  ▪ for s type: l=0; ml must be 0
	  ▪ for p type; l=1; ml can be -1, 0, or +1
	  ▪ for d type; l=2; ml can be -2, -1, 0, +1, or +2
	• ms (spin quantum number, represents the direction of spin of the electron)
	  ◦ can be +½ OR -½, not restricted by any other quantum number

Question Part 1 What is the maximum number of electrons contained in an orbital of type (x)? Of type (y)? Of type (z)? **Answer Part 1** type (x) = 2 electrons type (y) = 2 electrons type (z) = 2 electrons	Think about how many electrons can you put into any orbital? Think about electron **pairing**. Regardless of orbital type, each orbital can contain at most 2 electrons; one electron spins in a direction (m_s=+½), and the other electron spins in the opposite direction (m_s=-½).

	Guided Solution	Hint
tion	**Question Part 2** How many orbitals of type (x) are found in a shell with n = 2? How many of type (y)? How many of type (z)? First, identify each orbital type based on the shape in the provided figure:	Here we must consider what quantum numbers are allowed for each orbital type when n=2 type (x) - 1 lobe, sphere shaped (s); l=0 type (y) - 2 lobes, propeller shaped (p); l=1 type (z) - 4 lobes, daisy shaped (d); l=2
tion	Consider the relationship between quantum numbers n and l when n=2; is each orbital type (i.e., x, y, and z) allowed?	Recall: l must be zero or a positive whole number integer, restricted by n (i.e., l = 0, 1, 2, ... n-1) type (x) - when n=2; l=0 (allowed) type (y) - when n=2; l=1 (allowed) type (z) - when n=2; l≠2 (NOT allowed) Summary: a type (z) or 'd-type' orbital does not exist in the n=2 shell
tion	Consider the relationship between quantum numbers l and ml How many orientations of each allowed orbital exist when n=2?	Recall: for a given l value, the allowed values for ml are the total number of orbitals of that type that can exist type (x) - when n=2; l=0; ml = 0 • one allowed ml value therefore one 's' orbital when n=2 type (y) - when n=2; l=1; ml = -1, 0, +1 • three allowed ml values therefore three 'p' orbitals when n=2
tion	**Answer Part 2** Type (x) = 1 orbital Type (y) = 3 orbitals Type (z) = 0 orbitals	This makes sense because: an s-orbital (sphere shaped electron density cloud) is an identical sphere regardless of its orientation in space... resulting in only 1 possible s-orbital at each energy level a p-orbital (propellor shaped electron density cloud) can exist in three different axes in 3-dimensional space; it can lay along the x-axis, the y-axis, and the z-axis... resulting in 3 possible p-orbitals ($p_{x,}$ $p_{y,\ and}$ p_z) at each energy level
tion	**Question Part 3** Write a set of quantum numbers for an electron in an orbital of type (x) in a shell with n = 4. Of an orbital of type (y) in a shell with n = 2. Of an orbital of type (z) in a shell with n = 3. Consider what each quantum number represents? Summarize the quantum numbers you have been provided in the question.	Recall there are four distinct quantum numbers [n, l, ml ,ms] The question provides the principal quantum number (n) and the shape of the orbital defines the angular momentum quantum number (l): Type (x): [n=4, l=0] Type (y): [n=2, l=1] Type (z): [n=3, l=2]
tion	What quantum numbers must be defined to answer the question? Are there any restrictions that must be considered? Is there more than one correct answer?	Still need ml and ms • ml must be a positive or negative whole number integer, or zero; restricted by l (i.e., ml = -l, ... -2, -1, 0, 1, 2, ... +l) • ms can be +½ OR -½, not restricted by any other quantum number Yes, more than one allowed quantum number sets for each orbital type as defined in the question

	Guided Solution	Hint

tion	**Answer Part 3**
	type (x) = [4, 0, 0, ½]
	type (y) = [2, 1, 0, -½]
	type (z) = [3, 2, -1, ½]
		type (x) = [n=4, l=0, ml=0, ms=+½]
		• ml must equal 0 because l=0
		• ms = -½ would also be allowed
		  ○ 2 correct answers
		  ○ This makes sense because the 4s orbital can hold 2 electrons
		type (y) = [n=2, l=1, ml=0, ms=-½]
		• ml = -1 and ml = +1 would also be allowed
		• ms = +½ would also be allowed
		  ○ 6 correct answers
		  ○ This makes sense because the 2p orbitals can hold 6 electrons
		  ▪ 2 electrons in each p_x, p_y, and p_z
		type (z) = [n=3, l=2, ml=-1, ms=+½]
		• ml = -2, 0, +1, and +2 would also be allowed
		• ms = -½ would also be allowed
		  ○ 10 correct answers
		  ○ This makes sense because the 3d orbitals can hold 10 electrons
		  ▪ 2 electrons in each d_{xy}, d_{yz}, d_{xz}, $d_{x^2-y^2}$ and d_{z^2}

tion	**Question Part 4** What is the smallest possible n value for an orbital of type (x)? Of type (y)? Of type (z)? Consider the relationship between quantum numbers n and l.	Recall: l must be zero or a positive whole number integer, the value of l is restricted by n (i.e., l = 0, 1, 2, ... n-1; l≠n) Therefore the lowest value of n = (l+1)
tion	Recall the value of l for each type of orbital	type (x) - 1 lobe, sphere shaped (s); l=0 type (y) - 2 lobes, propeller shaped (p); l=1 type (z) - 4 lobes, daisy shaped (d); l=2
tion	**Answer Part 4** Type (x): n = 1 Type (y): n = 2 Type (z): n = 3	Type (x): lowest n =(l + 1)=(0 + 1)= 1 Type (y): lowest n =(l + 1)=(1 + 1)= 2 Type (z): lowest n =(l + 1)=(2 + 1)= 3
tion	**Question Part 5** What are the possible l and ml values for an orbital of type (x)? Of type (y)? Of type (z)? Consider the relationship between ml and l quantum numbers.	ml must be a positive or negative whole number integer, or zero; ml is restricted by l (i.e., ml = -l, ... -2, -1, 0, 1, 2, ... +l)
tion	Recall the value of l for each type of orbital	type (x) - 1 lobe, sphere shaped (s); l=0 type (y) - 2 lobes, propeller shaped (p); l=1 type (z) - 4 lobes, daisy shaped (d); l=2
tion	**Answer Part 5** Type (x): l=0; ml must be 0 Type (y): l=1; ml can be -1, 0, or +1 Type (z): l=2; ml can be -2, -1, 0, +1, or +2	Type (x) = s orbital: l=0; ml must be 0 Type (y) = p orbital; l=1; ml can be -1, 0, or +1 Type (z) = d orbital; l=2; ml can be -2, -1, 0, +1, or +2

(question source adapted question 6.3.11 from 6.E: Electronic Structure and Periodic Properties (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/06%3A_Electronic_Structure_and_Periodic_Properties_of_Elements/6.E%3A_Electronic_Structure_and_Periodic_Properties_(Exercises) , shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source question 41

https://openstax.org/books/chemistry-2e/pages/6-exercises , Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

2.3: Question 2.E.40 PASS - quantum numbers, orbital characteristics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.4: Question 2.E.58 PASS - electron configuration, orbital diagrams, Aufbau's Principle Hund's Rule

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 2.4.2.E.58

Consider magnesium (Mg), silicon (Si), and sulfur (S).

Which of the following atoms has two unpaired electrons?

a) Mg

b) Si

c) S

d) Both Mg and S

e) Both Si and S

Answer

e) Both Si and S

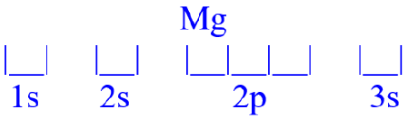
> See LibreText 2.6 Electronic Structure of Atoms

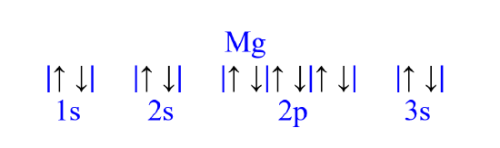
Strategy Map

Step	Hint
1. Find your element on the Periodic Table and note its atomic number	The atomic number (Z) describes the number of electrons the atom has Recall (LibreText section 3.1): The Periodic Table
2. Create an orbital box diagram for each elements electrons	Orbital diagrams are pictorial representations of the electron configuration, showing the individual orbitals as boxes and the pairing arrangement of electrons as arrows Begin to fill your box orbital diagram starting with the lowest energy '1s' orbital and stop when you “run out” of electrons • Follow Hund's Rule and Aufbau Principle
3. Count the number of unpaired electrons	Any orbitals containing one single electron are described as unpaired electrons

Solution

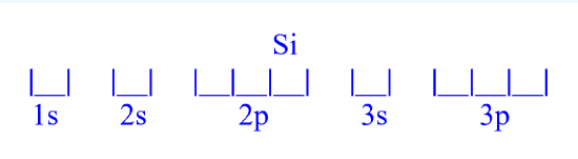
Magnesium (Mg): 12 electrons

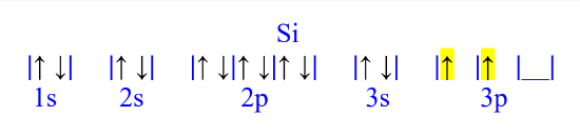




Mg has no unpaired electrons

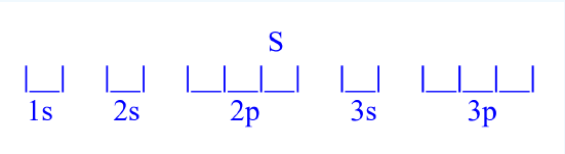
Silicon (Si): 14 electrons

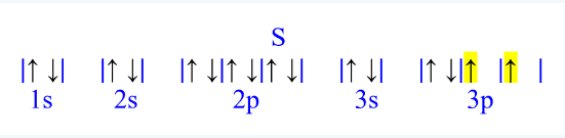




Silicon has two unpaired electrons

Sulfur (S): 16 electrons





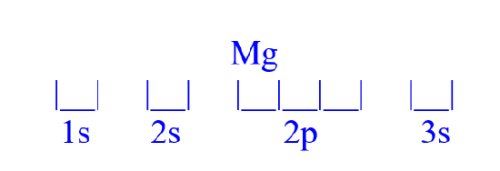
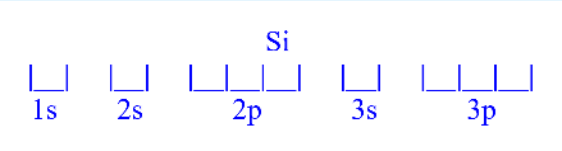
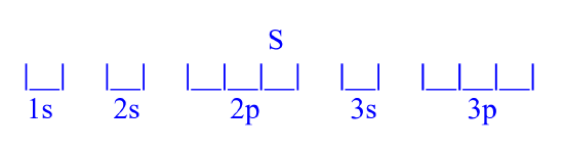
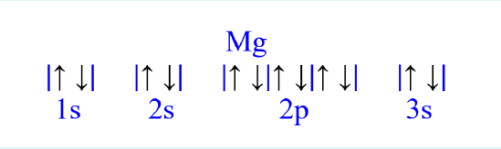
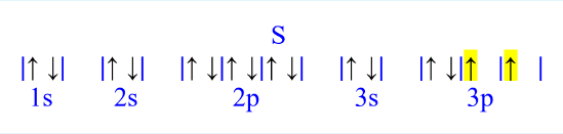
Sulfur has 2 unpaired electrons

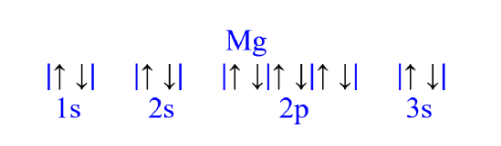
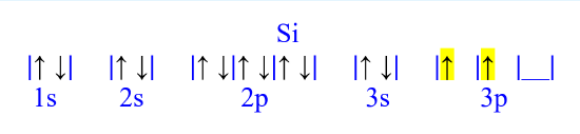
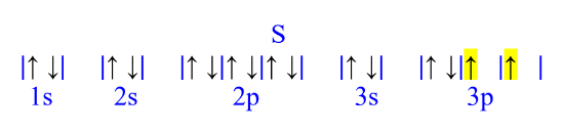
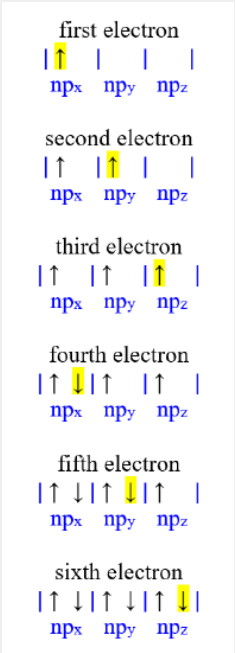
Answer

e) Both Si and S

Guided Solution

Guided Solution	Hint
This problem relies on concepts from Electronic Structure of Atoms (Electron Configurations)	Refer to LibreText 2.6, Electronic Structure of Atoms (Electron Configurations)
What does the question ask you to find? **Question** Consider magnesium (Mg), silicon (Si), and sulfur (S). Which of the following atoms has two **unpaired** electrons? Consider the circumstances in which electrons are paired or unpaired	• Aufbau Principle: *Fill electrons into the lowest available energy orbital* ◦ Paired electrons: If the lowest energy orbital available already has one electron, you will add a second electron to it with an opposite spin ▪ represented by an arrow pointed in the opposite direction • Hund's Rule: *when degenerate (same energy) orbitals are available, maximize the number of unpaired electrons with the same spin* ◦ Unpaired electrons: pay attention when filling the first, second, and third electrons into degenerate p-orbitals

	Guided Solution	Hint
tion	How many electrons does each element have? Mg: 12 electrons Si: 14 electrons S: 16 electrons	Find the atomic number (Z) on the Periodic Table to determine the number of electrons for each element
tion	Complete a Box Orbital diagram to determine the number of unpaired electrons for each element   	Common mistake: having an even number of electrons does not always mean that they are all paired! The orbitals in the box diagram are arranged from left to left to right in order of increasing energy • See LibreText Figure 2.6.2 - The Aufbau Principle
tion	Fill the appropriate number of electrons into each box diagram  Filled box orbital diagram for Si (1s2, 2s2, 2p6, 3s2, 3p2) 	Add electrons from left to right • maximize the number of electrons in each lowest energy available orbital before moving on ◦ each orbital can hold a maximum of two electrons ◦ the two electrons must have opposite spin (pointed up and down) • if degenerate empty orbitals are available, fill those first with unpaired electrons with the same spin before pairing electrons See LibreText Example 2.6.1: Quantum Numbers and Electron Configurations

Guided Solution	Hint
Complete Solution: **Mg**  Magnesium, atomic number 12, has a total of twelve electrons. The diagram is filled starting with a pair of electrons in the 1s orbital box, followed by 2s, the 2p then 3s. All electrons are paired in magnesium. **Si**  Silicon, atomic number 14, has a total of fourteen electrons. The diagram is filled starting with the 1s orbital box, followed by 2s, 2p, 3s then 3p. The 3p orbital has 2/6 electrons, meaning silicon has two unpaired electrons. **S**  Sulfur, atomic number 16, has a total of sixteen electrons. The diagram is filled starting with the 1s orbital box, followed by 2s, 2p, 3s then 3p. The 3p orbital has 4/6 electrons, meaning sulfur has two unpaired electrons.	When filling degenerate (same enegy) orbitals, Hund's Rule must be obeyed to maintain the most stable (lowest energy) electron configuration. Hund's Rule is shown below: 
Check your work! When completing the box orbital diagram, count the number of arrows; be sure it matches the number of electrons determined from the atomic number!	

(question source adapted question 6.4.14 from 6.E: Electronic Structure and Periodic Properties (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/06%3A_Electronic_Structure_and_Periodic_Properties_of_Elements/6.E%3A_Electronic_Structure_and_Periodic_Properties_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source question 59 https://openstax.org/books/chemistry-2e/pages/6-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

2.4: Question 2.E.58 PASS - electron configuration, orbital diagrams, Aufbau's Principle Hund's Rule is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

3: Periodic Relationships Among the Elements

Topic hierarchy

3.1: Question 3.E.05 PASS - ranking first ionization energy

3.2: Question 3.E.09 PASS - ranking atomic radius

3.3: Question 3.E.18 PASS - describe ionic radii of isoelectronic series

3: Periodic Relationships Among the Elements is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.1: Question 3.E.05 PASS - ranking first ionization energy

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 3.1.3.E.05

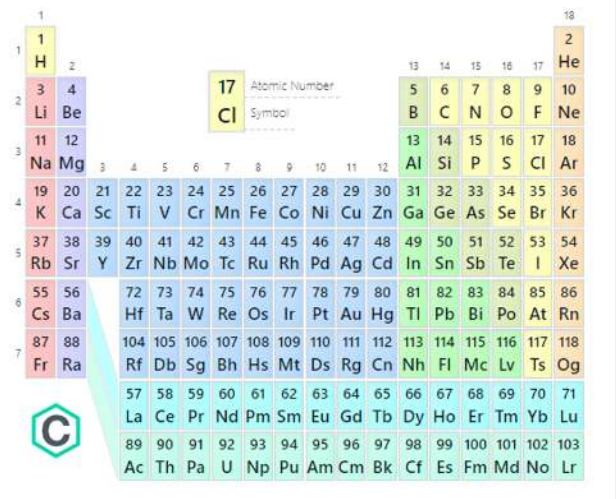
Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: F, Li, N, Rb

Answer

Rb, Li, N, F

> See LibreText 3.2 Periodic Variations in Elemental Properties

Strategy Map

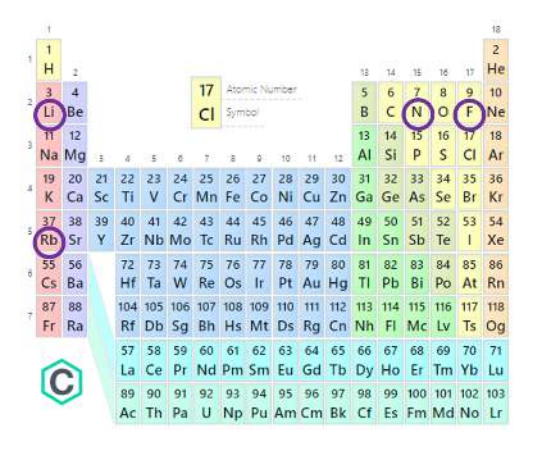
Step	Hint
1. Map out the order of the requested list	See LibreText 3.2 Periodic Variations in Elemental Properties Order the given elements by the amount of energy it would require removing their most loosely bound electron. • lowest/smallest to highest/greatest first ionization energy It must be ordered by the element that would require the least amount of energy to the element that would require the most. • less than < greater than
2. Recall the definition of ionization energy, what process does it describe?	Refer to LibreText Section 3.2.4: Variation in Ionization Energies
3. Summarize the trend in the periodic table	**Ionization Energies:** The ionization energies increase from left to right, and decrease from top to bottom.
5. Find the elements on the periodic table and arrange them based on the trend of increasing ionization energies.	

Solution

The locations of F, Li, N, Rb on the periodic table:

- F, atomic number 9, period 2, group 17
- Li, atomic number 3, period 2, group 1

- N, atomic number 7, period 2, group 15
- Rb, atomic number 37, period 5, group 1



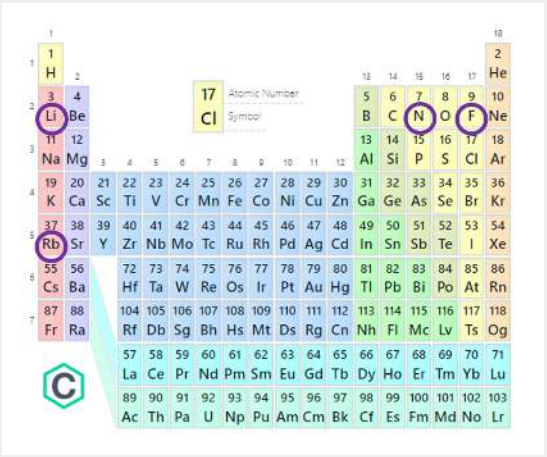
List smallest to largest first ionization energy: Rb < Li < N < F

Answer

In order of increasing first ionization energy: Rb, Li, N, F

Guided Solution

Guided Solution	Hint
This problem requires understanding of theory behind periodic trends Based on their positions in the periodic table, rank the following atoms in order of **increasing** first ionization energy: **F, Li, N, Rb**	See LibreText 3.2 Periodic Variations in Elemental Properties
This question asks us to rank the given elements by their ionization energies.	First ionization energy is "the amount of energy required to remove the most loosely bound electron from a gaseous atom in its ground state" Recall: "Energy is always required to remove electrons from atoms or ions, so ionization processes are endothermic and IE values are always positive."
Consider what characteristics may influence the amount of energy required to remove an electron from an atom • Effective Nuclear Charge (Zeff) • Atomic Radius	Recall that an electron is negatively charged, and the nucleus is positively charged. Consider what factors impact the strength of attractive (Coulombic) forces from the nucleus that are felt by the valence electron Refer to LibreText Section 3.2.1: Effective Nuclear Charge: Penetration and Shielding Refer to LibreText Section 3.2.2: Variation in Covalent Radius3.2: Periodic Variations in Element Properties
How does Zeff and atomic radius vary across the Periodic Table? Which characteristic is most influential on ionization energy?	**Effective Nuclear Charge (Zeff):** The effective nuclear charge increases from left to right and increases from top to bottom on the periodic table. **Atomic Radius:** The atomic radius decreases from left to right, and increases from top to bottom • Atomic radius has more impact on first ionization energy than effective nuclear charge

	Guided Solution	Hint
tion	How does first ionization energy vary across the Periodic Table?	**Ionization Energies:** The ionization energies increase from left to right, and decrease from top to bottom. • As atomic radius increases, first ionization energy decreases (less energy is required to remove the outer-most electron) • the smaller the atomic radius, the larger the first ionization energy (the more energy required to remove the outer-most electron)
tion	Locate the elements on the Periodic Table • F (fluorine), atomic number 9, period 2, group 17 • Li (lithium), atomic number 3, period 2, group 1 • N (nitrogen), atomic number 7, period 2, group 15 • Rb (rubidium), atomic number 37, period 5, group 1	
tion	List the elements in order of largest to smallest atomic radii Rb, Li, N, F	• (largest) Rb; due to principal quantum number n=5 > n=2 • Li > N > F (smallest); due to increasing effective nuclear charge (Zeff) moving left to right across the periodic table
tion	List smallest to largest first ionization energy: Rb, Li, N, F	smallest first ionization energy: Rb < Li < N < F largest first ionization energy
tion	**Does this make chemical sense?**	It is easiest to remove rubidium's valence electron because it is furthest from the nucleus and is held by the weakest attractive forces; it is hardest to remove fluorine's valence electron because it is closest to the nucleus and held by the strongest attractive forces

(question source 6.5.5 from 6.E: Electronic Structure and Periodic Properties (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/06%3A_Electronic_Structure_and_Periodic_Properties_of_Elements/6.E%3A_Electronic_Structure_and_Periodic_Properties_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source question 71 https://openstax.org/books/chemistry-2e/pages/6-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

3.1: Question 3.E.05 PASS - ranking first ionization energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.2: Question 3.E.09 PASS - ranking atomic radius

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 3.2.3.E.09

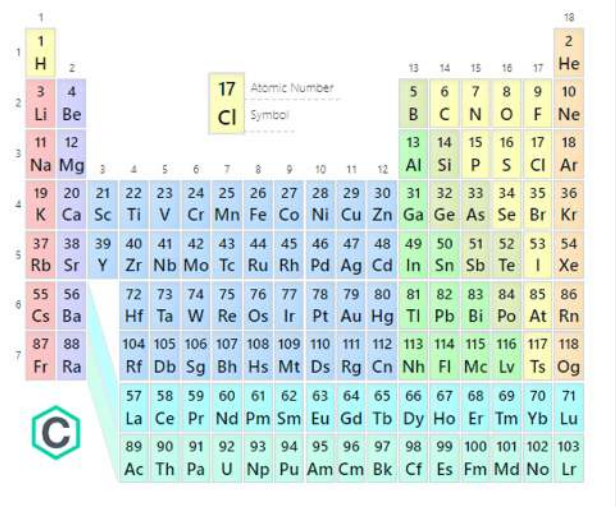
Based on their positions in the periodic table, list the following atoms in order of increasing radius: Ca, Cs, Mg, Rb.

Answer

Mg, Ca, Rb, Cs

> See LibreText 3.2 Periodic Variations in Elemental Properties

Strategy Map

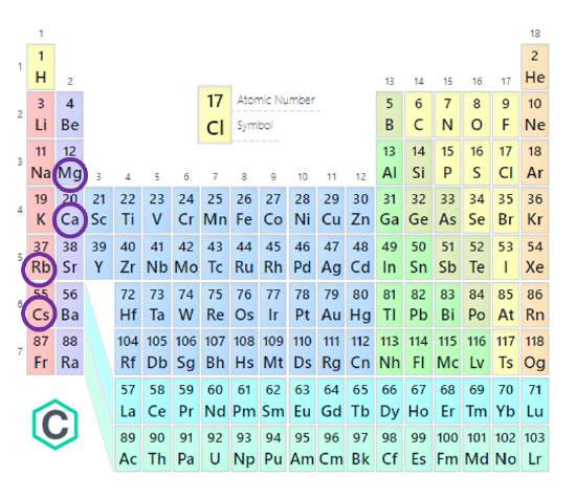
Step	Hint
1. Recognize the context of the question and confirm the requested format of the answer.	The question requires understanding of atomic radii Must list in order from smallest to largest atomic radii • Recall LibreText 3.2.2 Variation in Covalent Radius
2. Find each element on the periodic table.	Note the periods (row #; top to bottom), groups (column #; left to right) and atomic number (Z) 
3. Recall relevant concepts that impact atomic radii and summarize the periodic trends to sort them into descending or ascending sizes.	Consider effective nuclear charge (Zeff) • See LibreText 3.2.1 Effective Nuclear Charge: Penetration and Shielding Atomic radii increases as you move down a period, and decreases and you move right across and group • the largest element is located at the bottom left of the periodic table • the smallest element is located at the top right of the periodic table

Solution

The locations of Ca, Cs, Mg, and Rb on the periodic table:

- Ca, atomic number 20, group 2, period 4

- Cs, atomic number 55, group 1, period 6
- Mg, atomic number 12, group 2, period 3
- Rb, atomic number 37, group 1, period 5



List smallest to largest in atomic radii: Mg < Ca < Rb < Cs

Answer:

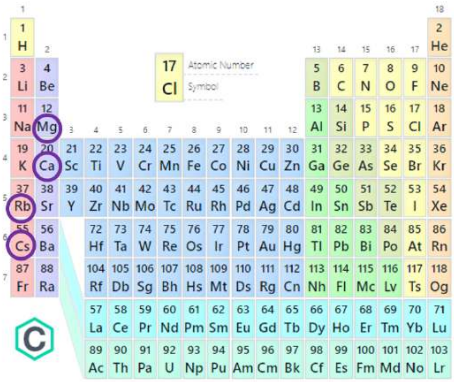
In order of increasing radius: Mg, Ca, Rb, Cs

Guided Solution

Guided Solution	Hint
This is a theory problem that requires understanding and application of concepts related to the shape of atomic orbitals, effective nuclear charge, and atomic radii.	See LibreText 2.5 The Shape of Atomic Orbitals • Recognize that as principal quantum number (n) increases (i.e., as you move down the periodic table) the size of the orbital increases See LibreText 3.2.1 Effective Nuclear Charge: Penetration and Shielding See LibreText 3.2.2 Variation in Covalent Radius • Recognize that as Zeff increases (i.e., as you move left to right across the periodic table) the size of the orbital decreases
Question Based on their positions in the periodic table, list the following atoms in order of **increasing radius**: Ca, Cs, Mg, Rb.	This question asks us to tank the given elements by their size. See LibreText Supplemental Modules (Physical and Theoretical Chemistry) - Atomic Radii

	Guided Solution	Hint

tion	Consider what characteristics may influence the distance between the nucleus and valence electron(s)	The nucleus is positively charged while electrons are negatively charged. The stronger the attractive force between the nucleus and the electron, the more tightly (i.e., closely) it will be held
		• Principal quantum number (n)
		  ○ Shielding: consider how many complete shells of electrons are between the valence electrons and the nucleus
		• Effective Nuclear Charge (Zeff)
		  ○ Shielding: electrons in the same shell (i.e., across the same row) are not effective at shielding the attractive forces from the nucleus
		  ○ as you move left to right across a row, the number of protons and effective nuclear charge (i.e., positive charge felt by the valence electron) increases
		  ○ the electrons are held more tightly (i.e., closer) to the nucleus and atomic radii decreases
		• (n) influences size more than (Zeff)

tion	Summarize the trend of atomic radii across the periodic table	The atomic radius decreases from left to right, and increases from top to bottom • the largest element is located at the bottom left of the periodic table • the smallest element is located at the top right of the periodic table
tion	Locate the elements on the Periodic Table • Ca (calcium), atomic number 20, group 2, period 4 • Cs (cesium), atomic number 55, group 1, period 6 • Mg (magnesium), atomic number 12, group 2, period 3 • Rb (rubidium), atomic number 37, group 1, period 5	

tion	List the elements in order of increasing principal quantum number (n)
	• Mg (period 3)
	• Ca (period 4)
	• Rb (period 5)
	• Cs (period 6)
		As (n) increases, the size of the orbital increases
		• Recall: The valence electrons for all of the listed elements are in s-orbitals
		  ○ Mg (3s)
		  ○ Ca (4s)
		  ○ Rb (5s)
		  ○ Cs (6s)
		• Note: Because all the elements are in different periods, we do not need to consider their group location (i.e., relative position from left to right)
		See Figure 3.2.3 in: LibreText 3.2.2 Variation in Covalent Radius

Guided Solution	Hint
Answer Mg, Ca, Rb, Cs	Listed from smallest to largest atomic radii: Mg < Ca < Rb < Cs
Does this make chemical sense?	It makes sense that size increases as principal quantum number increases. The periodic table is arranged logically with each row representing an increased (n). As the 'n' value increases, the size of the orbital is increased. This is because there are complete shells (i.e., core electrons) in between the nucleus and the valence electrons. The core electrons effectively shield the positive attractive force that is felt by the valence electrons. Therefore, the more core electrons, the weaker the attractive force; the valence electrons are held more weakly (i.e., at a further distance) from the nucleus resulting in an increased atomic radii.

(question source adapted from 6.5.9 from 6.E: Electronic Structure and Periodic Properties (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/06%3A_Electronic_Structure_and_Periodic_Properties_of_Elements/6.E%3A_Electronic_Structure_and_Periodic_Properties_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source question 75 https://openstax.org/books/chemistry-2e/pages/6-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

3.2: Question 3.E.09 PASS - ranking atomic radius is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.3: Question 3.E.18 PASS - describe ionic radii of isoelectronic series

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 3.3.3.E.18

The ionic radii of the ions S^{2-}, Cl^{-} and K^{+} are 184, 181, 138 pm respectively. Explain why these ions have different sizes even though they contain the same number of electrons.

Answer

Even though S^{2-}, Cl^{-} and K^{+} have the same number of electrons, the number of protons they have differs. This causes a difference in their effective nuclear charge (Zeff) and a variation in their size. The most negatively charged atom in an isoelectric series will have the greatest ionic radii due to the weakest Zeff and greatest electron repulsions.

> See LibreText 3.2 Periodic Variations in Elemental Properties
>
> See LibreText 3.2.3 Variation in Ionic Radii

Strategy Map

Step	Hint
1. Identify the difference between a parent atom and an ion	See LibreText 3.2.3 Variation in Ionic Radii Parent atom: atoms in their original state before any electrons are added or removed Ions: charged atoms that have either gained or lost electrons
2. Identify the change occurring when an element becomes a cation or anion	Anion: atom has gained electrons, is negatively charged • Ionization energy is related to the amount of energy required to remove an electron from an atom Cation: atom has lost electrons, is positively charged • Electron affinity is related to the amount of energy required to add an electron to an atom
3. Summarize the total number of protons and electrons for the ion	To find the number of protons and electrons for the neutral (uncharged) parent atom, refer to the periodic table atomic number (z). Note: a neutral atom has the same number of electrons and protons
4. Recall the relationship between protons and electrons, and consider the factors that influence their attraction	Protons are positively charged, electrons are negatively charged; they are attracted to each other by electrostatic 'Coloumbic' forces (just like oppositely charged magnets). • protons are contained within the nucleus, the centre of the atom • if there are more protons then electrons (positively charged), the electrons will be held more tightly (i.e., closer) to the nucleus, the ionic radii will be smaller • if there are more electrons than protons (negatively charged), the electrons will be held less tightly (i.e., further) from the nucleus, the ionic radii will be larger

Solution

Even though S^{2-}, Cl^- and K^+ have the same number of electrons, the number of protons they have differs. This causes a difference in their effective nuclear charge and a variation in their size.

The charge on an atom or ion is the sum of all electrons (-1) and protons (+1)

The parent atom has no charge, it has the same number of protons and electrons

- A cation (positively charged) always has **fewer electrons** and the same number of protons as the parent atom
  - will be **smaller** than the parent atom.
- An anion (negatively charged) always has **more electrons** and the same number of protons as the parent atom
  - will be **larger** than the parent atom.

In general, for an isoelectric series the more negatively charged an atom the larger it will be and the more positively charged an atom, the smaller it will be.

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory question requiring an explanation on why K^+, Cl^- and S^{2-} are not the same size even though they are an isoelectric series	Recall 'isoelectric series' is a group of elements (atoms and or/ions) that have the same number of electrons (i.e., the same electronic configuration) See LibreText 3.2 Periodic Variations in Elemental Properties See LibreText 3.2.3 Variation in Ionic Radii
tion	**Question** The ionic radii of the **ions S^{2-}, Cl^-, and K^+** are **184, 181, 138 pm** respectively. Explain why these ions have **different sizes** even though they contain the same number of electrons.	We are told the ions are varied sizes, we can use our knowledge to determine what an ion is and what the differences are between them Recall: when an ion is formed, the number of electrons changes, the number of protons remains the same.

tion	Consider the characteristics anions.	Recall:
		anions have gained electrons; electrons are negatively charged
		• There are more electrons (-1) than protons (+1)
		  ◦ the anion will be negatively charged
		this means two things:
		1. The relative effective nuclear charge (Zeff) from the nucleus is less
		  ◦ the valence electrons will be held less tightly (i.e., further away from the nucleus)
		2. The negative-negative repulsions between valence electrons is increased
		  ◦ when electrons are repelled by each other they move further apart
		Therefore: the ionic radii of anions is greater than the atomic radii of the parent atom
		See LibreText 3.2.1 Effective Nuclear Charge: Penetration and Shielding

Guided Solution	Hint

tion Consider the characteristics of cations	Recall:
	cations have lost electrons; electrons are negatively charged
	• There are less electrons (-1) than protons (+1)
	  ◦ the cation will be positively charged
	this means two things:
	1. The relative effective nuclear charge (Zeff) from the nucleus is greater
	  ◦ the valence electrons will be held more tightly (i.e., closer to the nucleus)
	2. The negative-negative repulsions between valence electrons is decreased
	Therefore: the ionic radii of cations is less than the atomic radii of the parent atom
	See LibreText 3.2.3 Variation in Ionic Radii

tion Summarize the total number of protons and electrons for the ion	Species / Atomic Number (Z) = Number of protons / Number of Electrons / Ion Charge **S^{2-}** / +16 / -18 / -2 **Cl^{-}** / +17 / -18 / -1 **K^{+}** / +19 / -18 / +1 The most negatively charged atom in an isoelectric series will have the greatest ionic radii due to the weakest Zeff and greatest electron repulsions.
tion **Answer** Even though S^{2-}, Cl^{-} and K^{+} have the same number of electrons, the number of protons they have differs. This causes a difference in their effective nuclear charge and a variation in their size. • K^{+} is the smallest because it has the most positive ionic charge • S^{2-} is the largest because it has the most negative ionic charge	

Guided Solution	Hint
Check your work! **Why does this answer make chemical sense?**	A cation is an atom that has lost at least one electron therefore is left with an unequal number of protons and electrons. Because there is a greater number of protons, the valence electron(s) will feel a greater effective nuclear charge and as a result be pulled closer to the nucleus in the center. This causes the cation to be smaller than its parent atom. The opposite is true for an anion, it has gained at least one electron and therefore has more electrons than it does protons. Because there are fewer protons then electrons, the valence electrons will feel a weaker effective nuclear charge. Additionally, the ionic radii will be affected by greater repulsions between the valence electrons and they will be pushed further from the nucleus. This causes the anion to be larger than its parent atom.

(question source adapted from 6.5.18 from 6.E: Electronic Structure and Periodic Properties (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/06%3A_Electronic_Structure_and_Periodic_Properties_of_Elements/6.E%3A_Electronic_Structure_and_Periodic_Properties_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source question 84 https://openstax.org/books/chemistry-2e/pages/6-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

3.3: Question 3.E.18 PASS - describe ionic radii of isoelectronic series is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

4: Chemical Bonding I - Basic Concepts

4.1: Question 4.E.20 PASS - which is the most polar bond?

4.2: Question 4.E.60 PASS - draw Lewis Structure and calculate formal charge

4.3: Question 4.E.65 PASS - determining enthalpy change from bond energies

PASS Lewis Structure H5P Activity Interactive Video (4.E.60)

4: Chemical Bonding I - Basic Concepts is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.1: Question 4.E.20 PASS - which is the most polar bond?

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 4.1.4.E.20

Which is the most polar bond?

C–C

C–H

N–H

O–H

Se–H

Answer

O-H is the most polar bond.

> See LibreText 4.3 Covalent Bonding (section 4.3.3 Electronegativity)

Strategy Map

Step	Hint
1. Assign your electronegativity values to each atom (use Pauling Electronegativity values from figure 4.3.3 of the LibreText).	Recall electronegativity represents the tendency of an atom to attract electron density (see LibreText section 4.3.3)
2. Identify the electronegativity difference between the elements forming the bond.	To do this, subtract the larger value from the smaller value.
3. Identify which bond has the largest electronegativity difference, that will be the most polar bond.	

Solution

C–C

Electronegativity difference = 0

C–H

C electronegativity = 2.5, H electronegativity = 2.1

Electronegativity difference = 0.4

N–H

N electronegativity = 3.0, H electronegativity = 2.1

Electronegativity difference = 0.9

O–H

O electronegativity = 3.5, H electronegativity = 2.1

Electronegativity difference = 1.4

Se–H

Se electronegativity = 2.4, H electronegativity = 2.1

Electronegativity difference = 0.3

O-H has the largest electronegativity difference so it is the most polar bond.

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory type problem that asks us to evaluate the covalent bonds between the elements given. We are asked to determine the most polar bond, which will be the bond with the most unequal electron distribution.	See LibreText 4.3 Covalent Bonding (section 4.3.3 Electronegativity)
tion	Which is the most polar bond? C–C C–H N–H O–H Se–H	
tion	The question asks you to identify which of the given bonds is the most polar, what property could relate to this?	Electronegativity is a measure of the tendency of an atom to attract electrons or electron density. Which bond has the highest electronegativity difference?
tion	Recall that polar covalent bonds have an electronegativity difference between the electron sharing atoms.	For this type of bond to occur the difference guideline is between 0.4 and 1.8. Within this range, the higher the difference, the more polar the bond. See LibreText 4.3 Covalent Bonding (section 4.3.5 Electronegativity and Bond Type)
tion	**Complete Solution:** **C–C** Same element type, no electronegativity difference so the bond is nonpolar. **C–H** C electronegativity = 2.5, H electronegativity = 2. Electronegativity difference = 0.4 **N–H** N electronegativity = 3.0, H electronegativity = 2.1 Electronegativity difference = 0.9 **O–H** O electronegativity = 3.5, H electronegativity = 2.1 Electronegativity difference = 1.4 **Se–H** Se electronegativity = 2.4, H electronegativity = 2.1 Electronegativity difference = 0.3 Difference less than 0.4 so the bond is nonpolar. The bond with the highest electronegativity difference is between O and H, therefore it is the most polar bond. **Answer: O-H**	For all bonds except C-C one of the atoms is H, so you can focus on the change in electronegativity of the atom H is bound to. If you rank these atoms from highest to lowest electronegativity you will predict which bond with H will be the most polar.

Guided Solution	Hint
Check your work! The most electronegative atom bound to H above is O with an electronegativity of 3.5, so we would predict that O-H would be the most polar bond in the list. **Why does this answer make chemical sense?** The more electronegative an atom is, the stronger the attraction between its nucleus and outside electrons will be. Atoms that are smaller will have higher electronegativities as there are fewer electrons to shield the nucleus from this attraction. Oxygen is one of the most highly electronegative elements as its nucleus forms a very strong attraction to outside electrons such as with a hydrogen atom.	Electronegativity ranking from largest to smallest: O > N > Se > C

(question source from page titled 7.E: Chemical Bonding and Molecular Geometry (Exercises) https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/07%3A_Chemical_Bonding_and_Molecular_Geometry/7.E%3A_Chemical_Bonding_and_Molecular_Geometry_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry-2e/pages/7-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

4.1: Question 4.E.20 PASS - which is the most polar bond? is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.2: Question 4.E.60 PASS - draw Lewis Structure and calculate formal charge

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 4.2.4.*E*.60

Iodine forms a series of fluorides (listed here). Draw Lewis structures for each of the four compounds and determine the formal charge of the iodine atom in each molecule:

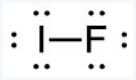
a) IF

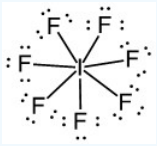
b) IF_3

c) IF_5

Answer

Lewis Structures:

a)  b) 

c)  d) 

Formal charges on iodine:

a) IF, formal charge on I = 0

b) IF_3, formal charge on I = 0

c) IF_5, formal charge on I = 0

d) IF_7, formal charge on I = 0

> See LibreText 4.4 Lewis Symbols and Structures and LibreText 4.5 Formal Charges and Resonance

Strategy Map

Step	Hint
1. Count valence electrons in the molecule.	Valence electrons are the outer shell electrons.
2. Draw the Lewis Structure: • Pick your center atom (if applicable) by selecting the atom that has the lowest electronegativity. • Arrange and connect your atoms. Start with a single bond represented by a single line between connected atoms. • Add lone pairs on terminal atoms until you have used them all. • Check that you followed the octet rule and that the total number of electrons you used (either as a dot/lone pair or in a line) add up to the total amount you counted in the beginning.	Recall your steps for drawing Lewis Structures (see LibreText section 4.4) If you have left over electrons and your center atom's octet is full, check and see if that center atom can have an expanded octet.

t va…

Step	Hint
1. Count valence electrons in the molecule.	Valence electrons are the outer shell electrons.
3. Calculate formal charge: • Identify the number of valence electrons the atom(s) of interest have. • Using your drawn Lewis Structure identify the number of valence electrons assigned to the atom(s) of interest. • Subtract your assigned electrons from the original valence electrons for each atom of interest. Charges can be positive or negative, zero means there is no formal charge for that atom.	Recall your steps to determine formal charge (see LibreText section 4.5) In a Lewis Structure, the assigned electrons are its surrounding lone pairs (one electron per dot, two electrons per pair) as well as half of the electrons from each line (this is because lines represent shared electrons, so each atom is assigned one from that pair).

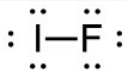
Solution (a) IF

Count valence electrons:

(1) Iodine + (1) Fluorine

$7ve^- + 7ve^- = 14$ valence electrons

Draw Lewis Structure:



Formal Charge on I:

Valence electrons	7
Assigned electrons	7
Formal Charge	0

Answer: The formal charge on I is 0.

Solution (b) IF_3

Count valence electrons:

(1) Iodine + (3) Fluorine

$7ve^- + 3(7ve^-) = 7 + 21 = 28$ valence electrons

Draw Lewis Structure:



Formal Charge on I:

Valence electrons	7
Assigned electrons	7
Formal Charge	0

Answer: The formal charge on I is 0.

Guided Solution (b) IF_3

Download Guided Solution as a pdf

	Guided Solution	Hint
tion…	This is a Lewis Structure application question. In this problem, you must follow the steps to create four different Lewis Structures and find their corresponding formal charges.	See LibreText 4.4 Lewis Symbols and Structures and LibreText 4.5 Formal Charges and Resonance
tion…	Iodine forms a series of fluorides (listed here). Draw Lewis structures for each of the four compounds and determine the formal charge of the iodine atom in each molecule: b) IF_3	
tion…	Count valence electrons in the molecule by identifying what atoms there are, counting the valence electrons each has and adding them.	To find how many valence electrons an atom has, count across the periodic table (skipping the D block elements). You can use the Group number or electron configuration to find the number of valence electrons in each type of atom. Group 1=1 Group 2=2 Group 13=3 Group 14=4 Group 15=5 Group 16=6 Group 17=7 Group 18=8 (full outer shell) You cannot have more electrons in a molecule than the amount each of the atoms had combined. For #1 Iodine has 7 valence electrons and Fluorine has 7 valence electrons, so the total number of electrons in the structure of IF will be 14.
tion…	Draw the Lewis Structure: • Pick your center atom (if applicable) by selecting the atom that has the lowest electronegativity. • Arrange and connect your atoms. Start by drawing all your atoms and connecting them with one line, this line represents a single bond and is your “bare-bones step” showing how the atoms are arranged. Each line represents two shared electrons. • Add lone pairs on terminal atoms until you have used them all, or until each atom has a full octet. • In the case that you run out of electrons before all octets are full, make double or triple bonds between sharing atoms. • In the case that you have remaining electrons, add them to the center atom until its octet is full. • Check that you followed the octet rule and that the total number of electrons you used (either as a dot/lone pair or in a line) add up to the total amount you counted in the beginning.	In our case **I** has a lower electronegativity, and **F** can only form one bond, so **I** will be the central atom and will be bonded to each **F**. Sometimes it can be confusing where to add a double bond. If you run out of available electrons, and some of your atoms still remain without a full octet, you can add a double bond between two or more atoms as long as it still abides by the octet rule. If you have left over electrons and your center atom's octet is full, check and see if that center atom can have an expanded octet.

Guided Solution	Hint
tion… Calculate formal charge: • Identify the number of valence electrons the atom(s) of interest have and add them up for each atom. • Using your drawn Lewis Structure identify the number of valence electrons **assigned** to the atom(s) of interest. • Subtract your assigned electrons from the original valence electrons for each atom of interest. Charge can be positive or negative, zero means there is no formal charge for that atom. **Element** **Element**	You counted these up to be able to draw your Lewis Strucrure. In a Lewis Structure, the assigned electrons are its surrounding lone pairs (one electron per dot, two electrons per pair) as well as half of the electrons from each line (this is because lines represent shared electrons, so each atom is assigned one from that pair). It can be useful to track valence electrons and calculate formal charges using a table as shown below, especially if you are working with many atoms: Element (Element name) (Element name) Valence electrons (ve^-) # # Assigned electrons (ae^-) # # Formal Charge (ve^- - ae^-) # #
tion… **Complete Solution b) IF_3** **Count valence electrons:** (1) Iodine + (3) Fluorine Iodine has 7 valence e^- Each Fluorine has 7 valence e^- and we have three F $7ve^- + 3x(7ve^-) = 28$ total valence electrons **Draw Lewis Structure:** Connect I and each F by a line representing a single bond. This has used 6 e-, and we have 22 left. Add lone pairs to each F atom. This adds 3 lone pairs to each F to complete their octet, for a total of 18 more electrons used. We have 4 electrons left. Add them as two lone pairs to the central I. This has used the remaining electrons.  **Determine Formal Charge:** Element Iodine Valence electrons 7 Assigned electrons 7 Formal Charge (7 - 7) 0 **Answer: The formal charge on I in IF_3 is 0.**	(Amount of Iodine atoms)x7Ve^- + (amount of fluorine atoms)x7ve^- = total valence electrons Each F atom has a complete octet so the octet rule is satisfied. I can have an expanded octet since it is in period 5. Looking back at the Lewis Structure assign one electron to I from the single bond, and count up all the electrons in the lone pairs on I.

Guided Solution	Hint
Check your work! The formal charge on a neutral molecule should always be zero. IF_3 is neutral so the zero formal charge is correct. **Why does this answer make chemical sense?** Lewis Structures are a visual representation of the electrons from an atom's outermost electron shell, to illustrate when they are shared in covalent bonds or present as lone pairs. Formal charges occur when the number of electrons an atom has is different from the amount it originated with prior to the reaction. A formal charge can be either positive or negative. Positive if there are less assigned than previously had, negative if more are assigned than previously had. This makes sense as electrons carry a negative charge.	Formal charges can also be used to help decide which of multiple structures is preferred. See LibreText 4.5 (section 4.5.2 Using Formal Charge to Predict Molecular Structure)

(question source from page titled 7.E:Chemical Bonding and Molecular Geometry (Exercises): https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/07%3A_Chemical_Bonding_and_Molecular_Geometry/7.E%3A_Chemical_Bonding_and_Molecular_Geometry_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry-2e/pages/7-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

link to PASS Lewis Structure H_5P Activity

4.2: Question 4.E.60 PASS - draw Lewis Structure and calculate formal charge is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.3: Question 4.E.65 PASS - determining enthalpy change from bond energies

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 4.3.4.E.65

Using the bond energies in Table 4.6.1, determine the approximate enthalpy change for each of the following reactions:

1. $\mathrm{H_2}(g) + \mathrm{Br_2}(g) \longrightarrow 2\,\mathrm{HBr}(g)$
2. $\mathrm{CH_4}(g) + \mathrm{I_2}(g) \longrightarrow \mathrm{CH_3I}(g) + \mathrm{HI}(g)$
3. $\mathrm{C_2H_4}(g) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{CO_2}(g) + 2\,\mathrm{H_2O}(g)$

Answer

1. −114 kJ
2. 30. kJ
3. −1055 kJ

> See LibreText 4.6: Strengths of Ionic and Covalent Bonds (section 4.6.1 Bond Strength: Covalent Bonds)

Strategy Map

	Step	Hint
	1. Identify which molecules have bonds that are being broken and which have bonds that are being formed.	Reactant bonds are broken, product bonds are formed in a chemical reaction.
ify …	2. Count up the bonds broken and formed. If you do not know what the bonds are create Lewis structures for each of the molecules.	Recall how to draw Lewis Structures (see LibreText section 4.4)
ify …	3. Use the provided table to look up the corresponding bond energies.	Table 4.6.1
ify …	4. Plug the values into the relationship that calculates enthalpy change from bond energies and solve.	To break bonds an energy input is required, when bonds formed energy is released. When energy is released it is negative. Recall the equation that demonstrates this relationship.

Solution

Reaction 1:

$\mathrm{H_2}(g) + \mathrm{Br_2}(g) \longrightarrow 2\,\mathrm{HBr}(g)$

$\Delta H = \sum \mathrm{D}_{\text{bonds broken}} - \sum \mathrm{D}_{\text{bonds formed}}$

$\Delta H = \sum \mathrm{D}_{\text{reactant bonds}} - \sum \mathrm{D}_{\text{product bonds}}$

$$\Delta H = (1mol(H-H) + 1mol(Br-Br)) - (2mol(H-Br))$$

$$\Delta H = (1mol(436kJ/mol) + 1mol(190.kJ/mol)) - (2mol(370.kJ/mol))$$

$$\Delta H = (626kJ) - (740.kJ) = -114kJ$$

Answer -114 kJ

Guided Solution Reaction 1

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a calculation type problem in which you must identify what bonds are being broken and formed during a chemical reaction. This type of problem requires you to use a table of values to calculate the overall energy change in the reaction using bond energies.	See LibreText 4.6: Strengths of Ionic and Covalent Bonds (section 4.6.1 Bond Strength: Covalent Bonds)
tion	Using the bond energies in Table 4.6.1, determine the approximate enthalpy change for each of the following reactions: 1. $H_2(g) + Br_2(g) \longrightarrow 2\,HBr(g)$	
tion	It can be difficult to know how many bonds are in a molecule. The easiest way to find what bonds there are is to create a Lewis Structure of the molecule.	Recall how to draw Lewis Structures (see LibreText section 4.4) H_2 Lewis Structure H–H Br_2 Lewis Structure :Br-Br: HBr Lewis Structure H-Br:
tion	When using the bond energies table, scan the table for your desired bond, i.e. (H-H), the number beside it will be the one used in the calculation. Make sure you don't confuse single, double, and triple bonds!	The bond energy for H-H is 436 kJ/mol.
tion	Pay attention to stoichiometry! Multiply the stoichiometric coefficient by your bond energy.	Students often make mistakes with this step. Ensure you are multiplying your bond energy value by the stoichiometric coefficient **as well as the number of times that bond shows up in the molecule**.
tion	Recall that it requires energy to break a bond (reactants) and energy is released when bonds are formed (products of a reaction). Remembering this should help you set up your equation to solve for ΔH!	$\Delta H = \sum D_{\text{bonds broken}} - \sum D_{\text{bonds formed}}$ $\Delta H = \sum D_{\text{reactant bonds}} - \sum D_{\text{product bonds}}$ The energy of bonds formed is subtracted as they are formed in the reaction, and bond energy values by definition are the energy required to break the bond. By subtracting the sum of the product bonds formed we are changing the sign of the energy to represent the opposite process (forming the bond).

	Guided Solution	Hint
tion	**Complete Solution:** The following equation will be used: $\Delta H = \sum \mathrm{D}_{\text{bonds broken}} - \sum \mathrm{D}_{\text{bonds formed}}$ For a reaction to occur, bonds need to break, and new bonds need to form. When bonds break energy is absorbed. When bonds are formed, energy is released. We use this to calculate the sum of the energy change over the reaction. How do we know which bonds are breaking and which are forming? **Identify and track your bonds:** $\mathrm{H}_2(g) + \mathrm{Br}_2(g) \longrightarrow 2\,\mathrm{HBr}(g)$ The bond between the two hydrogen atoms and the two bromine atoms are broken. **Two** product molecules are formed each with a hydrogen bonded to a bromine. $\Delta H = (1mol(H-H) + 1mol(Br-Br)) - (2mol(H-Br))$ **Look up the bond energies, enter them in equation:** $\Delta H = (1mol(436kJ/mol) + 1mol(190.kJ/mol)) - (2mol(370.kJ/mol))$ **Solve inside the brackets, remember stoichiometry:** $\Delta H = (626kJ) - (740.kJ)$ The energy released from the bonds formed are subtracted from the energy absorbed by the bonds broken. $\Delta H = (626kJ) - (740.kJ) = -114kJ$ The calculated value is negative, meaning the reaction was exothermic. **Answer** -114 kJ	In a reaction, the bonds to the left of the arrow will break and the bonds to the right of the arrow will form. ___(reactant bonds break)___ ⟶ ___(product bonds form)___ The bonds in these molecules are all single bonds. H-H 436 kJ/mol Br-Br 190. kJ/mol H-Br 370. kJ/mol Stoichiometric coefficients are exact numbers with infinite significant figures, so for H_2 is 1.000000..., so it is the bond energy values that determine significant figures for answer. You could also leave units of kJ/mol for the answer, since this represents kJ/mol of the reaction as written.
tion	**Check your work!** Comparing the bond energy values we see that the sum of the product bond (2 H-Br) have a greater magnitude than the sum of reactant bonds (H-H and Br-Br) so we would expect this reaction to have an exothermic ΔH. **Why does this answer make chemical sense?** When any reaction occurs, energy will be both absorbed and released, the overall sum however can be endothermic (requires more energy than it releases) or exothermic (it releases more energy than it requires). We can numerically observe this using the enthalpy change from bond energies equation.	Make sure you add up all bonds and consider the reaction stoichiometry.

(question source from page titled 7.E: Chemical Bonding and Molecular Geometry (Exercises) https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/07%3A_Chemical_Bonding_and_Molecular_Geometry/7.E%3A_Chemical_Bonding_and_Molecular_Geometry_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry-2e/pages/7-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

4.3: Question 4.E.65 PASS - determining enthalpy change from bond energies is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

PASS Lewis Structure H5P Activity Interactive Video (4.E.60)

PASS Lewis Structure H5P Activity Interactive Video (4.E.60) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

5: Chemical Bonding II - Molecular Geometry and Hybridization of Atomic Orbitals

5.1: Question 5.E.09 PASS - determine electron pair geometry and molecular geometry

5.2: Question 5.E.17 PASS - which central atom makes a polar molecule?

5.3: Question 5.E.59 PASS - identify central atom hybridization

5.4: Question 5.E.78 PASS - using molecular orbital theory and bond order to determine ion charge

---

5: Chemical Bonding II - Molecular Geometry and Hybridization of Atomic Orbitals is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.1: Question 5.E.09 PASS - determine electron pair geometry and molecular geometry

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 5.1.5.E.9

What are the electron-pair geometry and the molecular geometry of each of the following molecules or ions?

a) ClF_5

b) ClO_2^-

c) PCl_3

d) SeF_4

e) PH_2^-

Answer

a) Electron-pair geometry: octahedral, Molecular geometry: square pyramidal

b) Electron-pair geometry: tetrahedral, Molecular geometry: bent

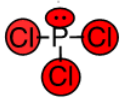
c) Electron-pair geometry: tetrahedral. Molecular geometry: trigonal bipyramidal

d) Electron-pair geometry: trigonal bipyramidal, Molecular geometry: seesaw

e) Electron-pair geometry: tetrahedral, Molecular geometry: bent

> See LibreText 5.1 Molecular Structure and Polarity (5.1.4 Molecular Polarity and Dipole Moment)

Strategy Map

Step	Hint
1. Create a Lewis Structure for the molecule.	Recall your steps for drawing Lewis Structures (see LibreText section 4.4): Arrange and connect your atoms. Add lone pairs on terminal atoms until you have used them all. Recall that Boron is an exception to the octet rule.
2. Identify how many total electron domains are attached to the center atom.	Electron domains are other atoms and lone pairs attached to the center atom. For example, PCl_3 has four electron domains. 
3. Identify how many of the electron domains are lone pairs. This will determine the molecules molecular geometry.	Lone pairs will repel the other atoms and distort the molecules shape. Example in PCl_3 the lone pairs attached to phosphorus cause the Chlorine atoms to be pushed downwards. 

Solution

a) ClF_5

42 valence electrons, Cl will be central atom.

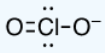


Six electron domains (The center Chlorine atom is attached to five Fluorine's and has one lone pair).

Electron-pair geometry: octahedral, Molecular geometry: square pyramidal

b) ClO_2^-

17 valence electrons, Cl will be central atom.

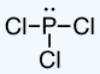


Four electron domains (The center Chlorine atom is attached to two oxygen's and has two lone pairs).

Electron-pair geometry: tetrahedral, Molecular geometry: bent

c) PCl_3

42 valence electrons, P will be central atom



Four electron domains (The center Phosphorus atom is attached to three Chorine's and has one lone pair).

Electron-pair geometry: tetrahedral. Molecular geometry: trigonal bipyramidal

d) SeF_4

44 valence electrons, Se will be central atom.



Five electron domains (The center Selenium is attached to four Fluorine's and has one lone pair).

Electron-pair geometry: trigonal bipyramidal, Molecular geometry: seesaw

e) PH_2^-

8 valence electrons, P will be central atom.



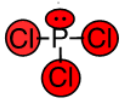
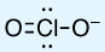
Four electron domains (The center phosphorus atom is attached to two hydrogens and has two lone pairs).

Electron-pair geometry: tetrahedral, Molecular geometry: bent

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory problem where you are asked to determine both the electron-pair geometries (The name for the shape given the number of domains) and the molecular geometries (The name of the shape depending on the electron cloud distortions) for each of the given molecules.	See LibreText 5.1 Molecular Structure and Polarity
What are the electron-pair geometry and the molecular geometry of each of the following molecules or ions?	

	Guided Solution	Hint
tion	To Figure out how many electron domains are in the molecule we must first draw the molecule with all lone pairs and bonds.	We can do this easily by first creating the molecules Lewis structure. For example, this shows that the molecule has four electron domains: 
tion	Why is the electron-pair geometry different from the molecular geometry?	Recall that the electron-pair geometry is the basic shape with a certain number of electron domains. The molecular geometry is the shape the molecule displays of the bonding domains, which is due to **electron pair repulsions. If there are lone pairs** that attached to the central atom they influence the shape of the molecule. The molecular geometry can be the same as the electron-pair geometry if the central atom **does not** have any **lone pairs**. For example: 
tion	**Complete Solution:** a) ClF_5 **Count valence electrons** 42 valence electrons **Connect your atoms** Cl will be central atom. **Draw the Lewis Structure**  **Count electron domains and bonding domains off the central atom** Six electron domains (The center Chlorine atom is attached to five Fluorine's and has one lone pair). Electron-pair geometry: **octahedral** Determine the molecular geometry by considering the resulting shape of the bonding domains only, and name the resulting molecular geometry.  Molecular geometry: **square pyramidal** b) ClO_2^- 17 valence electrons, Cl will be central atom.  Four electron domains (The center Chlorine atom is attached to two oxygen's and has two lone pairs). Electron-pair geometry: **tetrahedral**  Molecular geometry: **bent**	Valence electrons are the outer shell electrons. Cl is the least electronegative atom, so it is central. Total electron domains will determine electron pair geometry. Valence shell electron-pair repulsion theory (VSEPR theory) tells us that the electron pairs will arrange themselves to minimize repulsions by maximizing the distance between pairs. The molecular geometry or shape describes the relative positions of the bonds around a central atom. A wedge indicates a bond coming out of the plane of the screen, a dashed line indicates a bond going in to the plane of the screen. Add in one electron for the negative charge when counting valence electrons. O has a higher electronegativity so Cl will be the central atom. Add in one electron for the negative charge when counting valence electrons.

Guided Solution	Hint
c) PCl_3 42 valence electrons, P will be central atom.  Four electron domains (The center Phosphorus atom is attached to three Chorine's and has one lone pair). Electron-pair geometry: **tetrahedral**  Molecular geometry: **trigonal bipyramidal** d) SeF_4 44 valence electrons, Se will be central atom.  Five electron domains (The center Selenium is attached to four Fluorine's and has one lone pair). Electron-pair geometry: **trigonal bipyramidal**  Molecular geometry: **seesaw** e) PH_2^- 8 valence electrons, P will be central atom.  Four electron domains (The center phosphorus atom is attached to two hydrogens and has two lone pairs). Electron-pair geometry: **tetrahedral**  Molecular geometry: **bent**	
Check your work! Each of the above have lone pairs on the central atoms so we expect that the molecular geometry will be different than the electron pair geometry in each case and that is what our answers tell us. **Why does this answer make chemical sense?** Previously we had been looking at molecules two dimensionally as if they were flat like a paper. With molecular geometries we are beginning to see that molecules and ions have a three-dimensional shape that move freely. They will move the positions that require the least amount of energy. This means that the outside atoms will move the farthest they can from other negative charges (other atoms and electrons). Lone paired electrons are very repulsive and thus will cause even more distortion within an atom giving it a three-dimensional shape.	In order to determine electron pair geometry you must first start with a correct Lewis Structure. The bonding and lone pairs on the central atom are the electron domains. Electron pair geometry considers the three dimensional shape of all electron pairs. Molecular geometry considers the three dimensional shape of the bonding pairs.

(question source modified from page titled 4.11: Exercises https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_-_Atoms_First_2e_(OpenStax)/04%3A_Chemical_Bonding_and_Molecular_Geometry/4.11%3A_Exercises, shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry/pages/7-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

5.1: Question 5.E.09 PASS - determine electron pair geometry and molecular geometry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.2: Question 5.E.17 PASS - which central atom makes a polar molecule?

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 5.2.5.E.17

Molecule XF_3 has a dipole moment. Is X Boron or Phosphorus?

Answer

X is Phosphorus.

> See LibreText 5.1 Molecular Structure and Polarity (5.1.4 Molecular Polarity and Dipole Moment)

Strategy Map

Step	Hint
1. Add up the valence electrons for each molecule (BF_3 and PF_3).	
2. Draw the Lewis Structure for each molecule.	Recall your steps for drawing Lewis Structures (see LibreText section 4.4): • Arrange and connect your atoms • Add lone pairs on terminal atoms until you have used them all • Recall that Boron is an exception to the octet rule
3. Evaluate bond dipole moments and determine the shape of each molecule.	Differences in electronegativities between atoms will cause a bond to be polar. Lone pairs on the central atom influence molecular shape. When looking at the shape of a molecule, if the bond dipole vectors cancel out the molecule will be nonpolar. If the bond dipole vectors do not cancel out the molecule will have a net diploe moment and be polar.
4. Identify any differences between the two molecules that could cause one to be polar.	Which molecule has a lone pair on the central atom? What are the shapes of the two molecules?

Solution

Boron as X:

BF_3, 24 valence electrons, B will be central atom

Lewis structure 

Molecular Geometry, trigonal planar 

This molecule is non-polar.

Phosphorus as X:

PF_3, 26 valence electrons, P will be central atom

Lewis structure 

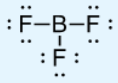
Molecular Geometry, trigonal pyramid 

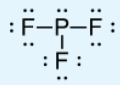
This molecule is Polar.

Since the molecule XF_3 has a dipole moment, X is Phosphorus.

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory type problem that tests you on your ability to identify polarity within molecules by comparing one polar and one non-polar molecule.	See LibreText 5.1 Molecular Structure and Polarity
tion	Molecule XF_3 has a dipole moment. Is X Boron or Phosphorus?	
tion	How does polarity occur within a molecule?	Differences in electronegativity between atoms will cause a bond to be polar. When looking at the shape of a molecule, if the bond dipole vectors cancel out the molecule will be nonpolar. If the bond dipole vectors do not cancel out the molecule will have a dipole moment and be polar.
tion	How do lone pairs on the central atom impact polarity?	Lone pairs on the central atom influence the molecular shape. The resulting positions of terminal atoms must be considered to see if the bond dipole moments are able to cancel
tion	**Complete Solution:** Boron as X, BF_3 **Count valence electrons** B has 3 valence electrons, each F has 7, total = 3 + 3(7) = 24 valence electrons **Connect your atoms** B is the central atom, each F is connected to B **Draw the Lewis Structure** Connect the B and each F by a single bond. Then place the remaining electrons as lone pairs around the terminal F atoms. Doing this uses 24 electrons, which matches our count of valence electrons, so the Lewis Structure is complete.  **Determine the molecular geometry or shape, and decide if the molecule is polar or nonpolar** Since there are three single bonds of the central atom and no lone pairs, the single bonds arrange themselves equidistant apart in a trigonal planar shape. The bond dipole vectors will all cancel and the molecule will be nonpolar.  trigonal planar, nonpolar Phosphorus as X, PF_3	Valence electrons are the outer shell electrons. The atom that can form the most bonds is the central atom, and is usually the least electronegative atom. Recall that Boron has a full octet with only six electrons, this is an exception to the octet rule. Valence shell electron-pair repulsion theory (VSEPR theory) tells us that the electron pairs will arrange themselves to minimize repulsions by maximizing the distance between pairs. The molecular geometry or shape describes the relative positions of the bonds around a central atom. Consider the bond dipoles in three dimensions to determine if they cancel.

Guided Solution	Hint
Count valence electrons P has 5 valence electrons, each F has 7, total = 5 + 3(7) = 26 valence electrons **Connect your atoms** P is the central atom, each F is connected to P **Draw the Lewis Structure** Connect the P and each F by a single bond. Then add electrons as lone pairs around the terminal F atoms. Then count how many valence electrons you have used. At this point we have used 24 valence electrons, and we have two left. Add a lone pair to P, which completes its octet. The Lewis Structure is complete.  **Determine the molecular geometry or shape, and decide if the molecule is polar or nonpolar** The lone pair and the three single bonds give a tetrahedral electron pair geometry. Since one is a lone pair the molecular geometry or shape of the molecule which results will be a trigonal pyramid. With this molecular geometry the bond dipole vectors cannot cancel out and the molecule will be polar.  trigonal pyramid, polar **Answer the question** Since the molecule XF_3 has a diploe moment, the molecule must be PF_3.	
Check your work! BF_3 is a trigonal planar molecule with no lone pairs and identical terminal atoms so it will be nonpolar. PF_3 has one lone pair on the central atom, has a tetrahedral electron pair geometry and a trigonal pyramid shape so it will be polar. **Why does this answer make chemical sense?** For a molecule to be non-polar, the net dipole moment must be zero, meaning the bond dipole moments cancel each other out when the molecular shape is considered. In the case where atom "X" is Boron there are no lone pairs on the central atom, the three bond dipole moments are equal in magnitude and in opposing directions when molecular shape is considered, the bond vectors cancel and the molecule is nonpolar. Since Phosphorus has a lone pair on the central atom, even though the bond dipole moments are equal they do not cancel each other out. This means the entire molecule has a dipole-moment and is polar.	If there are no lone pairs on the central atom and the terminal atoms are identical, the molecule will be nonpolar since the bond vectors cancel out. If there is a lone pair on the central atom and the bond vectors cannot cancel out the molecule will be polar.

(question source from page titled 4.11: Exercises https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_-_Atoms_First_2e_(OpenStax)/04%3A_Chemical_Bonding_and_Molecular_Geometry/4.11%3A_Exercises, shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source

https://openstax.org/books/chemistry/pages/7-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

5.2: Question 5.E.17 PASS - which central atom makes a polar molecule? is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.3: Question 5.E.59 PASS - identify central atom hybridization

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 5.3.5.E.59

Identify the hybridization of the central atom in each of the following molecules and ions that contain multiple bonds: (parts b-f)

(b) CS_2

(c) Cl_2CO (C is the central atom)

(d) Cl_2SO (S is the central atom)

(e) SO_2F_2 (S is the central atom)

(f) XeO_2F_2 (Xe is the central atom)

Answer

(b) sp

(c) sp_2

(d) sp^3

(e) sp3

(f) sp^3d

> See LibreText 5.3 Hybrid Atomic Orbitals

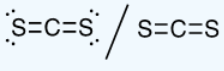
Strategy Map

Step	Hint
1. Create a Lewis Structure of the molecule.	See LibreText section 4.4 (Lewis Symblos and Structures)
2. Determine the Electron-Pair geometry of the molecule.	Recall using VSEPR to determine electron pair geometry (see LibreText section 5.1 Molecular Structure and Polarity)
3. Identify the hybridization of the center atom based on the molecule's Electron-Pair geometry.	The Hybridization of the center atom will be determined by the electron-pair geometry, not the atom itself. For example, in a linear molecule, the center atoms hybridization will always be sp.

Solution

(b) CS_2

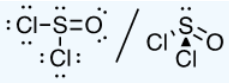
16 valence electrons, C central atom



This molecule has a linear electron-pair geometry: **sp hybridized**

(c) Cl_2CO (C is the central atom)

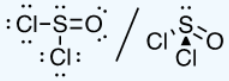
26 valence electrons



This molecule has a trigonal planar electron-pair geometry: **sp^2 hybridized**

(d) Cl_2SO (S is the central atom)

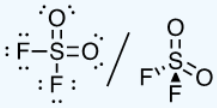
26 valence electrons



This molecule has a tetrahedral electron-pair geometry: **sp^3 hybridized**

(e) SO_2F_2 (S is the central atom)

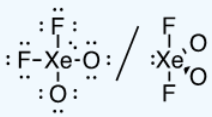
32 valence electrons



This molecule has a tetrahedral electron-pair geometry: **sp^3 hybridized**

(f) XeO_2F_2 (Xe is the central atom)

34 valence electrons

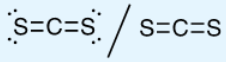
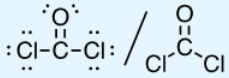
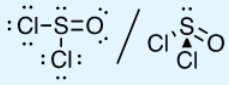
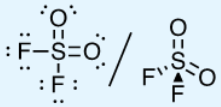
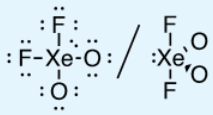


This molecule has a trigonal bipyramidal electron-pair geometry: **sp^3d hybridized**

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory problem that requires you to use your knowledge of molecular bonding and identify atomic orbital hybridization.	See LibreText 5.3 Hybrid Atomic Orbitals
tion	Identify the hybridization of the central atom in each of the following molecules and ions that contain multiple bonds: (only b-f) (b) CS_2 (c) Cl_2CO (C is the central atom) (d) Cl_2SO (S is the central atom) (e) SO_2F_2 (S is the central atom) (f) XeO_2F_2 (Xe is the central atom)	
tion	What is the pattern associated with atomic orbital hybridization?	The Hybridization of the center atom will be determined by the electron-pair geometry, not the atom itself. For example, in a linear molecule, the center atoms hybridization will always be sp. Number of hybrid orbitals = Number of orbitals mixed. Hybrids named by type and number of orbitals mixed.

Guided Solution	Hint
Complete Solution: (b) CS_2 16 valence electrons, C central atom  This molecule has a linear electron-pair geometry: **sp hybridized** (c) Cl_2CO (C is the central atom) 26 valence electrons  This molecule has a trigonal planar electron-pair geometry: **sp^2 hybridized** (d) Cl_2SO (S is the central atom) 26 valence electrons  This molecule has a tetrahedral electron-pair geometry: **sp^3 hybridized** (e) SO_2F_2 (S is the central atom) 32 valence electrons  This molecule has a tetrahedral electron-pair geometry: **sp^3 hybridized** (f) XeO_2F_2 (Xe is the central atom) 34 valence electrons  This molecule has a trigonal bipyramidal electron-pair geometry: **sp^3d hybridized**	Recognize this is a linear molecule, two electron domains: s + p = sp Hybridized The double bond counts as one electron domain. three electron domains: s + p + p = sp^2 hybridized The double bond counts as one electron domain, the lone pair is one electron domain, and the two single bonds are each one electron domain. four electron domains: s + p + p + p = sp^3 hybridized four electron domains: s + p + p = sp^3 hybridized five electron domains: s + P + p + d = sp^3d hybridized

Guided Solution	Hint
Check your work! Verify your Lewis Structure account for all the valence electrons and that you have counted electron domains off the central atom correctly. The name of our electron pair geometry and number of electron domains leads us to the correct hybridization. **Why does this answer make chemical sense?** Hybrid orbitals form due to the covalent bonding within molecules. When two or more atoms create a covalent bond, their atomic orbitals will overlap and form such hybrid orbitals. In this case, we are looking at the hybridization of the center atom to see what kind of hybridization is happening to overlap with all the atoms. When molecules get larger and more atoms are involved, more bonds will be made, and thus more orbitals will be involved in bonding. This is why a linear molecule only uses two orbitals (s + p) and an octahedral molecule uses six orbitals (s + p + p + p + d + d).	Valence electrons are outer shell electrons. Look at the condensed electron configuration.

(question source from page titled 8E: Advanced Theories of Covalent Bonding (Execricse) ttps://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/08%3A_Advanced_Theories_of_Covalent_Bonding/8.E%3A_Advanced_Theories_of_Covalent_Bonding_(Exercises) - Chemistry LibreTexts, shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, https://openstax.org/books/chemistry/pages/8-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

5.3: Question 5.E.59 PASS - identify central atom hybridization is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.4: Question 5.E.78 PASS - using molecular orbital theory and bond order to determine ion charge

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 5.4.5.E.78

What charge would be needed on F_2 to generate an ion with a bond order of 2?

Answer

The charge would be positive 2, or 2+.

> See LibreText 5.5 Molecular Orbital Theory

Strategy Map

Step	Hint
1. Fill out a molecular orbital bond energy diagram.	Recall molecular orbital energy diagrams (see LibreText section 5.5.1)
ut a … 2. Calculate the bond order of F_2.	Recall bond order calculations (see LibreText section 5.5.2)
ut a … 3. Calculate the number of electrons it would require for a bond order of 2.	
ut a … 4. Identify the electron difference and if you would have to add or subtract electrons to F_2 to fix it.	

Solution

Bond Order formula:

$$\text{bond order} = \frac{(\text{number of bonding electrons}) - (\text{number of antibonding electrons})}{2}$$

The typical bond order for F_2:

$$\text{bond order} = \frac{\text{Bonding electrons} - \text{antibonding electrons}}{2}$$

$$\text{bond order} = \frac{8-6}{2}$$

$$\text{bond order} = 1$$

Electrons required for bond order of 2:

$$\text{bond order} = \frac{\text{Bonding electrons} - \text{antibonding electrons}}{2}$$

$$2 = \frac{x}{2}$$

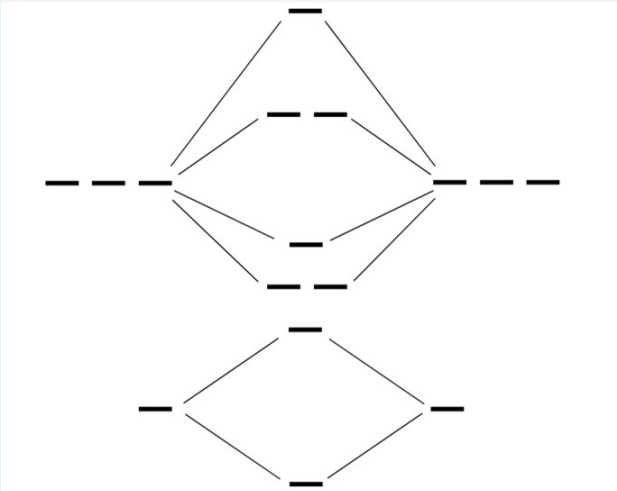
$$X = 4$$

Two electrons would have to be removed, therefore the charge on the molecule would be 2+.

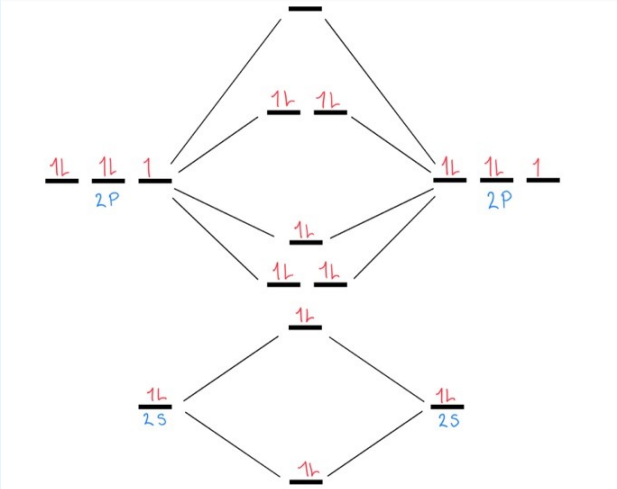
Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint

	Guided Solution	Hint
tion	This is both a theory and calculation type problem where you will use the bond order equation to find the charge a polyatomic ion would require having a bond order of 2.	
tion	What charge would be needed on F_2 to generate an ion with a bond order of 2?	See LibreText 5.5 Molecular Orbital Theory
tion	Try filling out a molecular orbital energy diagram for F_2.  This shows how many bonding and antibonding electrons are in this polyatomic ion which you will use for your calculation.	Check your molecular orbital energy diagram for F_2 by clicking here
tion	Recall the bond order equation.	$\text{bond order} = \frac{(\text{number of bonding electrons}) - (\text{number of antibonding electrons})}{2}$
tion	Since we are adding a charge to the polyatomic ion, we know that we are manipulating the number of electrons it has. This means you will have to identify the difference in electrons for the polyatomic ions normal bond order and if it were to have a bond order of two.	Use the bond order equation to calculate the number of electrons it would require having a bond order of two: $2 = \frac{x}{2}$ Subtract the number you get for "x" from the number of electrons F_2 would usually have.
tion	Will the charge be positive or negative?	Recall that when you add electrons to an atom it gains a negative charge and when you remove electrons its gains a positive charge.
tion	**Complete Solution:** Bond Order formula: $\text{bond order} = \frac{(\text{number of bonding electrons}) - (\text{number of antibonding electrons})}{2}$ The typical bond order for F_2: molecular orbital energy diagram	

Guided Solution



There are eight bonding electrons and six antibonding electrons in F_2.

$$\text{bond order} = \frac{\text{Bonding electrons} - \text{antibonding electrons}}{2}$$

$$\text{bond order} = \frac{8-6}{2}$$

$$\text{bond order} = 1$$

The bond order for a ground state F_2 polyatomic ion would be four.

Electrons required for bond order of 2:

Since we know we want the bond order to be two, but we wont know the number of electrons that would be, we can use the same formula to isolate and solve for the number of electrons.

$$\text{bond order} = \frac{\text{Bonding electrons} - \text{antibonding electrons}}{2}$$

$$2 = \frac{x}{2}$$

$$X = 4$$

If we remove two of the antibonding electrons from F_2 its new bond order would be two.

$$\text{bond order} = \frac{8-4}{2} = \frac{4}{2} = 2$$

Since two electrons are being removed, F_2 would gain a positive charge of 2 making the polyatomic ion F_2^{2+}.

Hint

Check your work!

For a bond order of 2, two electrons have been removed from F_2 leaving it with a charge of positive 2.

Why does this answer make chemical sense?

The bond order is the number of bonds in a molecule or polyatomic ion. The higher the bond order, the stronger the bond. It makes sense for the bond order to increase if two electrons are removed from F_2, the single bonded atom would have to become double bonded to fulfill the octet rule. A double bond is much stronger than a single bond, this can be mathematically seen when looking at the table of bond energies, it requires significantly more energy to break a double bond than a single bond.

(question source from page titled 8.E: Advanced Theories of Covalent Bonding (Exercises) https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/08%3A_Advanced_Theories_of_Covalent_Bonding/8.E%3A_Advanced_Theories_of_Covalent_Bonding_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry/pages/8-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

5.4: Question 5.E.78 PASS - using molecular orbital theory and bond order to determine ion charge is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

6: Intermolecular Forces and Liquids and Solids

6.1: Question 6.E.07 PASS - define and explain the types of intermolecular forces

6.2: Question 6.E.11 PASS - arrange in order of increasing boiling point

6.3: Question 6.E.15 PASS - predict melting point

6: Intermolecular Forces and Liquids and Solids is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.1: Question 6.E.07 PASS - define and explain the types of intermolecular forces

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 6.1.6.E.7

Define the following and give an example of a molecule in which each occurs as the strongest intermolecular attraction:

a. dispersion force

b. dipole-dipole attraction

c. hydrogen bond

Answer

a) Dispersion Force

A dispersion force is an intermolecular attraction between temporary dipoles which form due to asymmetric electron distribution.

Example: Br_2

b) Dipole-Dipole Attraction

A dipole-dipole attraction is an intermolecular attraction between two permanent dipoles, which results from an electrostatic attraction of the positive end of one polar molecule to the negative end of another polar molecule.

Example: ICI

c) Hydrogen Bonds

A hydrogen bond is the electrostatic attraction between the partially positive hydrogen atom (when the hydrogen is bound to F, O, or N) in one molecule and the partially negative atom (F, O, or N) in another molecule.

Example: HF

> See LibreText 6.1 Intermolecular Forces

Strategy Map

Step	Hint
1. Identify if polarity plays a role with that type of force, this will help determine its cause.	If polarity does not play a role, the type of interaction is spontaneous and non-permanent. If polarity does play a role, the type of interaction is due to the molecule having partial charges on either end.
2. Identify the strength of the interaction.	There are different types of dipole-dipole attractions. The strongest type of diploe-dipole attraction has its own name.
3. Recall which molecules the interactions occur in.	

ify i…

ify i…

Solution

a) Dispersion Force

A dispersion force is an intermolecular attraction between temporary dipoles which form due to asymmetric electron distribution.

Dispersion forces occur as an atom develops a temporary dipole moment when its electrons are distributed asymmetrically about the nucleus, and is then attracted to another atom with a temporary diploe moment. They are the intermolecular

attraction which occurs between nonpolar molecules.

Example: Br_2

b) Dipole-Dipole Attraction

A dipole-dipole attraction is an intermolecular attraction between two permanent dipoles, which results from an electrostatic attraction of the positive end of one polar molecule to the negative end of another polar molecule.

For example the positive end of one polar molecule will be attracted to the negative end of another polar molecule.

Example: ICI

c) Hydrogen Bonds

A hydrogen bond is the electrostatic attraction between the partially positive hydrogen atom (when the hydrogen is bound to F, O, or N) in one molecule and the partially negative atom (F, O, or N) in another molecule.

The electrostatic attraction between the partially positive hydrogen atom in one molecule and the partially negative atom in another molecule gives rise to a very strong type of dipole-dipole interaction called a hydrogen bond.

Example: HF

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory type problem that asks you to define three different types of intermolecular forces and provide three molecules to exemplify each force.	
tion	Question: Define the following and give an example of a molecule in which each occurs as the strongest intermolecular attraction: a. dispersion force b. dipole-dipole attraction c. hydrogen bond	See LibreText 6.1 Intermolecular Forces
tion	Recall your knowledge of polarity.	Polarity occurs when a molecule has an unsymmetrical net dipole, it will cause one atom to gain a partial positive charge and another to gain a partial negative charge. If the molecule exhibits constant polarity, you can expect its intermolecular interactions to be a form of dipole-dipole attraction.
tion	Why is there an attraction between two molecules? (think about magnets) Are some attractions stronger than others?	Attraction between molecules occurs due to electrostatic attractions (when a negatively charged atom is attracted a positively charged atom). Intermolecular forces are weaker forces that occur between molecules. There are variations of strength between the types of intermolecular forces. Dispersion forces are the weakest, dipole-dipole attractions are stronger and hydrogen bonds are a very strong type of dipole-dipole attraction.
tion	**Complete Solution:** **a) Dispersion Force** A dispersion force is an intermolecular attraction between temporary dipoles which form due to asymmetric electron distribution.	See LibreText 6.1 Intermolecular Forces (section 6.1.2 Dispersion Forces) See LibreText 6.1 Intermolecular Forces (section 6.1.3 Dipole-Dipole Attractions) See LibreText 6.1 Intermolecular Forces (section 6.1.4 Hydrogen Bonding)

Guided Solution

Dispersion forces occur as an atom develops a temporary dipole moment when its electrons are distributed asymmetrically about the nucleus.

A second atom can then be distorted by the appearance of the dipole in the first atom. The electrons of the second atom are attracted toward the positive end of the first atom, which sets up a dipole in the second atom. The net result is rapidly fluctuating, temporary dipoles that attract and are caused by one another.

Dispersion forces can occur between all molecules. Electron clouds are constantly moving and attracting to other positive charges around them. For this reason, a momentary disturbance can cause a dispersion force between two or more molecules.

Dispersion forces occur as an atom develops a temporary dipole moment when its electrons are distributed asymmetrically about the nucleus, and is then attracted to another atom with a temporary diploe moment. They are the intermolecular attraction which occurs between nonpolar molecules.

Answer

A dispersion force is an intermolecular attraction between temporary dipoles which form due to asymmetric electron distribution.

Example: Br_2

b) Dipole-Dipole Attraction

A dipole-dipole attraction is an intermolecular attraction between two permanent dipoles, which results from an electrostatic attraction of the positive end of one polar molecule to the negative end of another polar molecule.

A dipole-dipole attraction is an intermolecular attraction between two permanent diploes, which results from an electrostatic attraction of the positive end of one polar molecule to the negative end of another polar molecule.

Polar molecules have partially charged atoms due to unsymmetrical net dipoles. This negative partial charge can easily form an attraction to another partial positive charged atom. For example the positive end of one polar molecule will be attracted to the negative end of another polar molecule.

Answer

A dipole-dipole attraction is a force that results from an electrostatic attraction of the positive end of one polar molecule to the negative end of another polar molecule.

Example: ICI

c) Hydrogen Bonds

A hydrogen bond is the electrostatic attraction between the partially positive hydrogen atom (when the hydrogen is bound to F, O, or N) in one molecule and the partially negative atom (F, O, or N) in another molecule.

Hydrogen bonds form whenever a hydrogen atom is bonded to one of the more electronegative atoms, such as a fluorine,

Hint

Guided Solution	Hint
oxygen, or nitrogen atom. The electrostatic attraction between the partially positive hydrogen atom in one molecule and the partially negative atom in another molecule gives rise to a very strong type of dipole-dipole interaction called a hydrogen bond. Polarity is caused by electronegativity differences. Hydrogen has a very low electronegative and forms a very strong bond when paired with high electronegative atoms. When electronegativity differences occur, the higher electronegative atom will pull the shared electrons closer to itself. In the case with hydrogen, the highly electronegative atom pulls the electrons closer to its nucleus and further from hydrogen's nucleus giving hydrogen a very strong positive charge. When two of these molecules come into contact, they will form a very strong attraction called a hydrogen bond. **Answer** A hydrogen bond is the electrostatic attraction between the partially positive hydrogen atom (when the hydrogen is bound to F, O, or N) in one molecule and the partially negative atom (F, O, or N) in another molecule. Example: HF	
Check your work! Look at the type of molecule you have listed as the example for each intermolecular attraction and make sure it fits and that it would have that attraction as its dominant force. For teh above examples Br_2 is nonpolar, ICl is a polar covalent molecule but cannot form hydrogen bonds, and HF is a polar covalent molecule that can form hydrogen bonds. **Why does this answer make chemical sense?** Intermolecular forces are forces that occur between molecules. All molecules have intermolecular forces at some point. The types of forces they are influenced by are dependent on the properties of the molecule. Dispersion forces are felt by all molecules as they are spontaneous and occur due to temporary charges of the moving electron clouds. Polar molecules will experience dipole-dipole attractions and the strength of the polarity will determine the strength of the interaction. Hydrogen bonds occur between very polar molecules where atoms have a high electronegativity difference. When electronegativity differences occur, the higher electronegative atom will pull the shared electrons closer to itself. In the case with hydrogen, the highly electronegative atom pulls the electrons closer to its nucleus and further from hydrogen's nucleus giving hydrogen a very strong positive charge. The attraction between two of these special molecules is the strongest of intermolecular attractions.	

(question source from page titled 10.E: Liquids and Solids (Exercises) https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/10%3A_Liquids_and_Solids/10.

E%3A_Liquids_and_Solids_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry/pages/7-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

6.1: Question 6.E.07 PASS - define and explain the types of intermolecular forces is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.2: Question 6.E.11 PASS - arrange in order of increasing boiling point

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 6.2.6.*E*.11

Arrange each of the following sets of compounds in order of increasing boiling point temperature:

a) HCl, H_2O, SiH_4

b) F_2, Cl_2, Br_2

c) CH_4, C_2H_6, C_3H_8

Answer

a) $SiH_4 < HCl < H_2O$

b) $F_2 < Cl_2 < Br_2$

c) $CH_4 < C_2H_6 < C_3H_8$

Strategy Map

Step	Hint
1. Identify what intermolecular forces are acting on each of the molecules in question.	See LibreText 6.1 Intermolecular Forces
ify … 2. Rank the molecules based on their intermolecular forces (the stronger the forces, the higher the boiling point)	Intermolecular forces ranked from strongest to weakest: Hydrogen Bonding > Dipole-Dipole Forces > Dispersion Forces
ify … 3. If two or more molecules have the same intermolecular forces, rank them based on the appropriate properties.	When comparing nonpolar molecules, the larger and heavier a molecule is the stronger the dispersion forces, and the higher the boiling point temperature.

Solution

a) HCl, H_2O, SiH_4

SiH_4 **dispersion forces**

HCl **dipole-dipole attractions** + dispersion forces.

H_2O **hydrogen bonding** + dispersion forces

Answer: $SiH_4 < HCl < H_2O$

b) F_2, Cl_2, Br_2

F_2 **dispersion forces**

Cl_2 **dispersion forces**

Br_2 **dispersion forces**

The largest molecule has the highest boiling point temperature.

Answer: $F_2 < Cl_2 < Br_2$

c) CH_4, C_2H_6, C_3H_8

All are nonpolar with dispersion forces.

The larger the molecule (the larger the surface area), the higher the boiling point temperature.

Answer: $CH_4 < C_2H_6 < C_3H_8$

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory type problem where you will rank the given compounds by their properties (boiling point). This can be done by evaluating the strength of their intermolecular forces.	
Arrange each of the following sets of compounds in order of increasing boiling point temperature: a) HCl, H_2O, SiH_4 b) F_2, Cl_2, Br_2 c) CH_4, C_2H_6, C_3H_8	
How is boiling point impacted by intermolecular forces?	The stronger the intermolecular forces, the more energy required to break the intermolecular forces and transition the compound between the liquid phase and the gas phase. The stronger the intermolecular forces are, the more energy it would require to boil.
Recall the comparative strength of the intermolecular forces.	Intermolecular forces ranked from strongest to weakest: Hydrogen Bonding > Dipole-Dipole Forces > Dispersion Forces
If the molecules are impacted by the same types of intermolecular forces, which one will have the higher boiling point temperature?	If the molecules in question all have the same intermolecular forces, their boiling point temperatures will be dependent on the chemical properties of the molecule. For nonpolar molecules the size of the molecule will change its boiling point, the larger the molecule (molar mass/atomic radius/surface area) the higher the boiling point temperature.

Guided Solution	Hint
Complete Solution: - Identify all forces acting in each molecule. - Bolded are the most dominant forces in the molecule (keeping in mind that hydrogen bonding is a type of dipole-dipole attraction). - Hydrogen bonding is the strongest force, followed by dipole-dipole attractions and the weakest force is dispersion forces. Dispersion forces act in all molecules, if they are the only forces acting between the molecules, a “tie breaker” is used to determine the boiling point temperature. a) HCl, H_2O, SiH_4 SiH_4 **dispersion forces** HCl **dipole-dipole attractions** + dispersion forces. H_2O **hydrogen bonding** + dispersion forces **Answer:** $SiH_4 < HCl < H_2O$ b) F_2, Cl_2, Br_2 F_2 **dispersion forces** Cl_2 **dispersion forces** Br_2 **dispersion forces** All forces are the same type therefore it goes based on size. The largest atom has the highest boiling point temperature. **Answer:** $F_2 < Cl_2 < Br_2$ c) CH_4, C_2H_6, C_3H_8 All are nonpolar with dispersion forces. All molecules have the same elements but differ in the numbers of C and H and the molecule size. The larger the molecule (larger size, molar mass and surface area), the higher the boiling point temperature. **Answer:** $CH_4 < C_2H_6 < C_3H_8$	

Guided Solution	Hint
Check your work! The question asked you to arrange in order of increasing boiling point temperature, so the lowest boiling point is first. The molecules will be arranged in order of weakest intermolecular attractions to the strongest (going from left to right) **Why does this answer make chemical sense?** All molecules have intermolecular forces which determine the phase of a substance. Solids have the most intermolecular forces, liquids have less and gases have the least. To transition between phases intermolecular forces must be formed or broken. A boiling point is the transition between a liquid state and a gaseous state which requires the breaking of intermolecular forces. Breaking intermolecular forces requires energy (heat) and the amount of energy is determined by the intermolecular forces themselves. The stronger the force, the more energy is required. Hydrogen bonds are the strongest due to them being between two very polar molecules. Dipole-Dipole attractions are slightly weaker, but still stronger than dispersion forces as the molecules in a dipole-dipole attraction both have a constant partial charge. Dispersion forces are the weakest and occur in all molecules due to the movement of electron clouds. Within dispersion forces, molecules with higher surface areas will have stronger attractions than smaller molecules due to there being more space for the molecules to “cling” to one another. (Think about Velcro strips, the larger strip will have a better grip than the smaller one due to its larger surface area).	'<' means 'less than' Make sure you are using this representation correctly.

(question source from page titled Liquids and Solids (Exercises) https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/10%3A_Liquids_and_Solids/10.E%3A_Liquids_and_Solids_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry-2e/pages/10-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

6.2: Question 6.E.11 PASS - arrange in order of increasing boiling point is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.3: Question 6.E.15 PASS - predict melting point

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 6.3.6.*E*.15

The melting point of $H_2O(s)$ is 0 °C. Would you expect the melting point of $H_2S(s)$ to be −85 °C, 0 °C, or 185 °C? Explain your answer.

Answer

−85 °C

Water has hydrogen bonds which are a stronger intermolecular force. H_2S must therefore melt at a lower temperature since it has diploe-dipole intermolecular attractions which are weaker than hydrogen bonds.

> See LibreText 6.1 Intermolecular Forces

Strategy Map

Step	Hint
1. Identify the difference between the two molecules.	Recall Lewis Structures of both molecules (see LibreText section 4.4)
2. Identify what intermolecular forces each molecule has.	They do not have the same intermolecular forces.

Solution

Answer: −85 °C.

Water has hydrogen bonds between the molecules which are a stronger intermolecular force, so it takes more energy to break some intermolecular attractions and change from a solid to a gas, so we expect H_2S to melt at a lower temperature than H_2O.

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory type problem where you must identify the intermolecular forces of the given molecules and decide if the melting point of one will be lower, the same, or higher than the other.	See LibreText 6.1 Intermolecular Forces
The melting point of $H_2O(s)$ is 0 °C. Would you expect the melting point of $H_2S(s)$ to be −85 °C, 0 °C, or 185 °C? Explain your answer.	
Hydrogen bonding will occur when a hydrogen atom is bonded to what other atoms?	Recall that hydrogen bonding occurs when a hydrogen atom is bonded to Fluorine, Oxygen, and Nitrogen.
Since the molecules have different atoms you can expect that the molecules **do not have** the **same** melting point.	The molecule with the stronger intermolecular forces will have the higher melting point.

	Guided Solution	Hint
tion	**Complete Solution:** **Identify the intermolecular attractions in each molecule** Although both H_2O and H_2S have hydrogen atoms, Sulfur is not one of the very electronegative atoms, therefore H_2S does not contain hydrogen bonds, it contains dipole-dipole attractions. **Compare the strength of the intermolecular attractions** Dipole-Dipole attractions are not as strong as hydrogen bonds meaning the melting point of H_2S must be lower than the melting point of H_2O. **Answer the question** The melting point of H_2S is expected to be **−85 °C.** Some intermolecular attractions must be broken to melt a substance in order to change it form a solid to a liquid, so molecules with stronger intermolecular attractions will melt a a higher temperature, molecules with weaker intermolecular attractions will melt at a lower temperature. Since H_2S has dipole-dipole intermolecular attractions which are weaker than the stronger hydrogen bonds that H_2O contains, H_2S must have a lower melting point than H_2O.	Students often see a hydrogen atom and think of a hydrogen bond, but they must remember that hydrogen bonds only occur between hydrogen and specific atoms.
tion	**Check your work!** We expect our H_2S melting point to be less than that for water, and −85 °C is less than the 0 ºC melting point of water. **Why does this answer make chemical sense?** The two molecules do not have the same intermolecular forces. Hydrogen bonding is caused by very strong electronegativity difference between hydrogen and one of the highly electronegative atoms. The strong difference between the two causes the hydrogen atom to have a very strong partial positive charge. Although the bond between hydrogen and sulfur is strong, it is only a dipole-dipole attraction, meaning the forces between two H_2O molecules is stronger than the forces between two H_2S molecules. Since H_2S has a weaker force, its melting point will require less energy and thus will be lower than H_2O.	

(question source from page titled Liquids and Solids (Exercises) https://chem.libretexts.org/Bookshelves/General_Chemistry/Chemistry_1e_(OpenSTAX)/10%3A_Liquids_and_Solids/10.E%3A_Liquids_and_Solids_(Exercises), shared under a CC BY 4.0 license, authored, remixed, and/or curated by OpenStax, original source https://openstax.org/books/chemistry-2e/pages/10-exercises, Access for free at https://openstax.org/books/chemistry/pages/1-introduction)

6.3: Question 6.E.15 PASS - predict melting point is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

7: Organic Chemistry I - Bonding and Structure

7.1: Question 7.1.7 PASS - label hybridized atomic orbitals, identify sigma and pi bonds

7.2: Question 7.2 PASS - interpret chemical structure line drawing

7.3: Question 7.4.6 PASS - identifying constitutional isomers

7.4: Question 7.5.2.1 PASS - naming cycloalkane structures

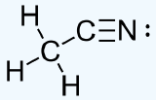
7: Organic Chemistry I - Bonding and Structure is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.1: Question 7.1.7 PASS - label hybridized atomic orbitals, identify sigma and pi bonds

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 7.1.7

For the molecule acetonitrile:



a) How many sigma and pi bonds does it have?

b) What orbitals overlap to form the C-H sigma bonds?

c) What orbitals overlap to form the C-C sigma bonds?

d) What orbitals overlap to form the C-N sigma bonds?

e) What orbitals overlap to form the C-N pi bonds?

f) What orbital contains the lone pair electrons on nitrogen?

g) Sketch the molecule showing the hybridized atomic orbital 'cloud' overlap.

Answer

a) 5 sigma and 2 pi.

b) A sp^3 hybrid orbital from carbon and an s orbital from hydrogen

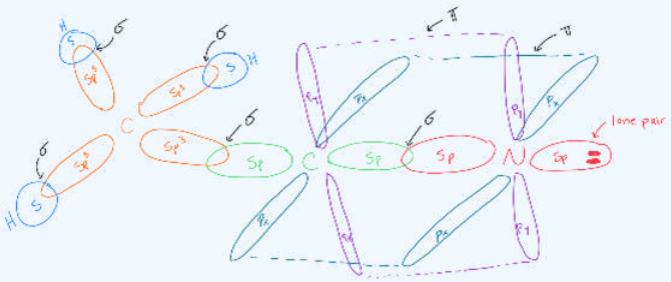
c) A sp^3 hybrid orbital from one carbon and a sp hybrid orbital from another carbon.

d) A sp hybrid orbital from carbon and a sp orbital from nitrogen.

e) A p_y and a p_z orbital from carbon and a p_y and p_z orbital from nitrogen.

f) A sp hybrid orbital.

g)



See LibreText 7.1.7 sp hybrid orbitals and the structure of acetylene

Strategy Map

Step	Hint

Step	Hint
1. Count the number of single bonds (sigma) in the molecule	There is a single bond between every single atom. Recall Lewis Structures (see LibreText section 4.4.2) and Valence Bond Theory (see LibreText section 5.2)
1. Count the number of single bonds (sigma) in the molecule	There is a single bond between every single atom. Recall Lewis Structures (see LibreText section 4.4.2) and Valence Bond Theory (see LibreText section 5.2)
2. Count the number of double and triple bonds (pi) in the molecule.	Double/triple bonds are represented as second and third lines along a single bond.
3. Identify hybridization of center atoms (the two carbon atoms).	The hybridization of the center atoms will depend on their electron domain geometry. Recall VSEPR Theory (see LibreText section 5.5.1)
4. Identify the hybridization of the outer atoms, if any.	
5. Identify which atoms are overlapping. The hybrid orbitals that overlap will depend on steps 3 and 4.	
6. Sketch your diagram showing all overlaps and ensure you label everything.	

Solution

a) 5 sigma (single bonds) and 2 pi (one double and one triple bond).

b) A sp^3 hybrid orbital from carbon and an s orbital from hydrogen. (The hydrogen atoms are all unhybridized.)

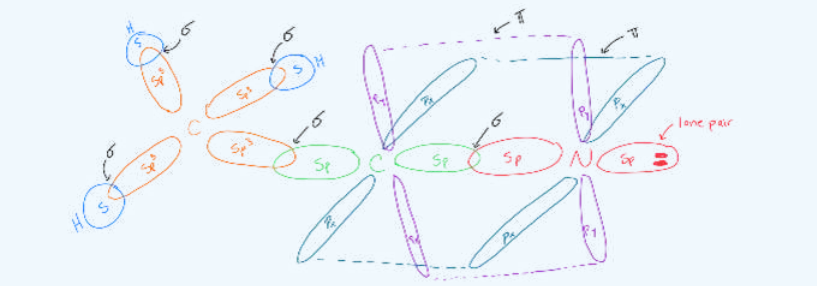
c) A sp^3 hybrid orbital from one carbon and a sp hybrid orbital from another carbon. (Both carbons are hybridized, but their electron domain geometries are not the same, meaning their hybridizations are different.)

d) A sp hybrid orbital from carbon and a sp orbital from nitrogen. (Both carbon and nitrogen have a sp hybridization.)

e) A p_y and a p_z orbital from carbon and a p_y and p_z orbital from nitrogen. (the 'Y' and 'Z' just mean they have slightly different orientations)

f) A sp hybrid orbital. (Nitrogen has a linear electron domain geometry meaning it has a sp hybridization in both directions.)

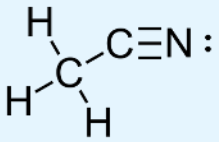
g) Hand drawn sketch:

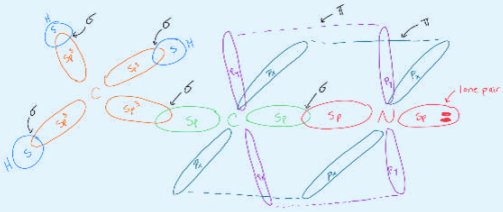
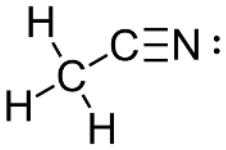


Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint

	Guided Solution	Hint
tion	This is a theory type problem that tests your knowledge on atomic orbital hybridization. You must be able to correctly identify the hybridization of multiple atoms in a molecule and sketch the orbital overlaps.	See LibreText Section 7.1.7 (sp Hybrid Orbitals and the Structure of Acetylene)
tion	For the molecule acetonitrile:  a) How many sigma and pi bonds does it have? b) What orbitals overlap to form the C-H sigma bonds? c) What orbitals overlap to form the C-C sigma bonds? d) What orbitals overlap to form the C-N sigma bonds? e) What orbitals overlap to form the C-N pi bonds? f) What orbital contains the lone pair electrons on nitrogen? g) Sketch the molecule showing the hybridized atomic orbital 'cloud' overlap.	
tion	What are sigma and pi bonds? Sigma bonds are your single bonds. Pi bonds are your double and triple bonds.	There is at least one sigma bond between each of the connected atoms. Look at the bonds between your carbon and your nitrogen. How many lines do you see??
tion	How do you identify atomic orbital hybridization? 1. Identify the electron domain geometry of your central atom. (in this case, the two center atoms are carbons) 2. Your electron domain geometry will tell you the number of bonding directions attached to your atom. 3. Each bonding direction represents an orbital needed. 4. Identify if your outer atoms require hybridization (if they do, they will have multiple bonding directions as well, this includes lone pairs)	For instance, a tetrahedral electron domain geometry must have four electron domains so it will have an s+p+p+p (sp^3) hybridization.
tion	How do you know which orbitals are overlapping?	Each bond that is between the atoms will be an overlap. The hybridization of the orbitals will depend on the two atoms that the bond is between.
tion	How do you sketch the orbital overlaps? 1. Begin by drawing your center atom with its orbitals in the shape of its electron domain geometry. (it's okay if they just look like "sausages" if you label them with their hybridization) 2. Overlap these orbitals with orbitals representing the outer atoms. 3. Label these as sigma bonds. 4. The orbitals that overlap to form pi bonds are longer and skinnier than the sigma bonds. They overlap sideways and therefore connect above and below the sigma bond.	Using different colours in your drawing for orbital type can be helpful to keep track.

Guided Solution	Hint
Complete Solution: a) There will always be a single bond between each atom in the molecule. A second or third line next to the single bond represents a double or triple bond. You can count these in your Lewis structure form to identify the sigma and pi bonds. **answer** 5 sigma (single bonds) and 2 pi (one double bond and one triple bond). b) This carbon atom makes bonds in four directions which gives it a sp^3 hybridization. Three of these directions are hydrogen atoms. All of these hydrogen atoms are unhybridized and have s orbitals. When the orbitals from the hydrogen and carbon overlap, they form a sigma bond. **answer** A sp^3 hybrid orbital from carbon and an s orbital from hydrogen. c) Although both atoms are hybridized, they have a different number of bonding directions therefore one is sp^3 and the other is sp. **answer** A sp^3 hybrid orbital from one carbon and a sp hybrid orbital from another carbon. d) Both atoms make bonds in two different directions giving them sp orbital hybridization. **answer** A sp hybrid orbital from carbon and a sp orbital from nitrogen. e) These are the pi bonds. They are labeled with 'Y' and 'Z' as they have slightly different orientations on the axis. Pi orbitals make bonds above and below the sigma bond as they overlap sideways. **answer** A p_y and a p_z orbital from carbon and a p_y and p_z orbital from nitrogen. f) Nitrogen has two electron domain directions. One is towards the carbon in a triple bond and the other is towards its lone pair therefore, It has two hybridized sp orbitals. The orbital that contains nitrogens lone pair is an sp orbital. **answer** A sp hybrid orbital. g) Hand drawn sketch below: **answer** 	

Guided Solution	Hint
Check your work! Go back over your answers with the sketch and the structure to make sure you have tracked all the bonds and orbitals correctly. There are sigma bonds between each of the connected atoms, and the two pi bonds are the second and third bond between the C and N. Make sure that your identified orbitals match the electron domain geometry. **Why does this answer make chemical sense?** Hybridized orbitals are created so that all single bonds are made by the same orbitals and are equal. This is why hybridization will depend on the number of bonding directions made by the atom. The hybridization will likely vary between the atoms in the bond but will remain consistent within an atom itself.	

(question source from page titled 1.9: sp Hybrid Orbitals and the Structure of Acetylene, shared under a CC BY-SA 4.0 license, authored, remixed, and/or curated by Steven Farmer, Dietmar Kennepohl, Krista Cunningham, Tim Soderberg, William Reusch, & William Reusch. Access for free at https://chem.libretexts.org/Bookshelves/Organic_Chemistry/Organic_Chemistry_(Morsch_et_al.)/01%3A_Structure_and_Bonding/1.09%3A_sp_Hybrid_Orbitals_and_the_Structure_of_Acetylene)

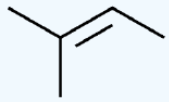
7.1: Question 7.1.7 PASS - label hybridized atomic orbitals, identify sigma and pi bonds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.2: Question 7.2 PASS - interpret chemical structure line drawing

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 7.2

How many carbons are in the following drawing? How many hydrogens?



Answer

The structure has five carbon atoms and ten hydrogen atoms.

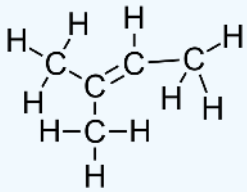
> See LibreText 7.2 Drawing Chemical Structures

Strategy Map

Step	Hint
1. Identify the carbons in the line drawing for the molecule.	In a line structure, carbon atoms are found at the end of each line as well as at each corner.
2. Add hydrogen atoms to the carbons until each carbon has 4 bonds.	This will satisfy the octet rule.
3. Count the carbon atoms and the hydrogen atoms in the molecule.	

Solution

Structure showing all bonds:



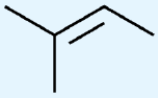
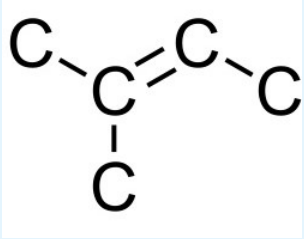
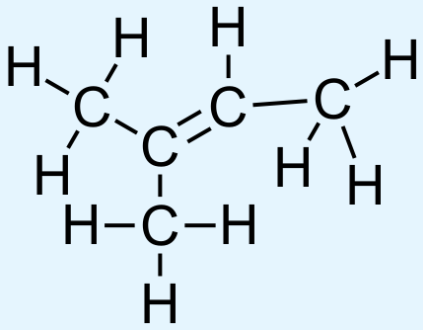
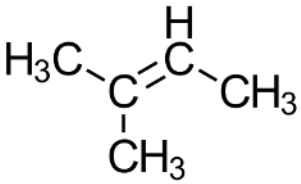
Answer:

The full Lewis Structure shows that in this molecule there are five carbon atoms and ten hydrogen atoms.

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory problem that requires you to identify all implied carbon and hydrogen atoms in a line structure.	Refer to LibreText 7.2 Drawing Chemical Structures

	Guided Solution	Hint
tion	Question: How many carbons are in the following drawing? How many hydrogens? 	
tion	What do the lines represent? Where are the implied carbon and hydrogen atoms?	Recall that in a line structure, carbon atoms are found at the end of each line as well as at each corner. Hydrogen atoms surround the carbons to satisfy the octet rule.
tion	Draw the structure of the molecule showing all carbon and hydrogen atoms. Remember to fulfill the octet rule.	Recall your steps for drawing Lewis Structures (see LibreText section 4.4)
tion	**Complete Solution:** In a line structure, carbons are at either end of each line as well as in every corner. A double line shows the double bond. Start by writing in the carbon atoms and their connectivity based on the line diagram:  Then add hydrogen atoms to each C atom until each C has 4 bonds:  Count the Carbon and Hydrogen atoms shown in the structure. **Answer** In this molecule there are five carbon atoms and ten hydrogen atoms	Recognize the impact of the double bond in the structure. You could also add the atoms to a condensed structure, just ensure the octet rule is satisfied on each carbon. 

Guided Solution	Hint
Check your work! Counting valence electrons, there are 4 for each C and 1 for each H making a total of 30 valence electrons. The Lewis Structure shows 30 valence electrons and satisfies the octet rule. **Why does this answer make chemical sense?** The full structure represents the Lewis Structure for this molecule. There are 5 carbon atoms which matches the line structure given in the question, and 10 hydrogen atoms makes the octet rule satisfied for each carbon. There are many ways to draw molecules. A line structure gives us the ability to quickly create carbon-hydrogen structures. By convention we know that each line represents the carbon-to-carbon bond, and all hydrogen atoms are "hidden."	Recall your steps for drawing Lewis Structures (see LibreText section 4.4)

(question source from page titled 1.12: Drawing Chemical Structures https://chem.libretexts.org/Bookshelves/Organic_Chemistry/Organic_Chemistry_(Morsch_et_al.)/01%3A_Structure_and_Bonding/1.12%3A_Drawing_Chemical_Structures, shared under a CC BY-SA 4.0 license, authored, remixed, and/or curated by Steven Farmer, Dietmar Kennepohl, Krista Cunningham, Tim Soderberg, William Reusch, & William Reusch)

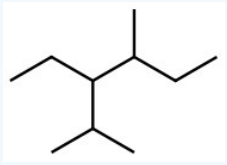
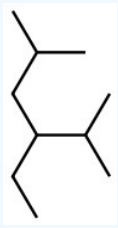
7.2: Question 7.2 PASS - interpret chemical structure line drawing is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

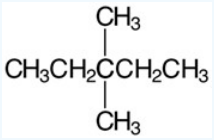
7.3: Question 7.4.6 PASS - identifying constitutional isomers

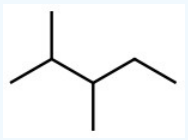
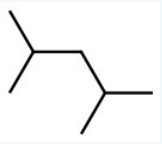
This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

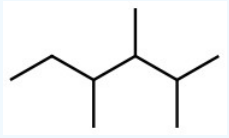
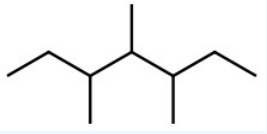
Exercise 7.3.7.4.6

Indicate whether each of the following sets are constitutional isomers, the same compound, or different compounds.

a)  & 

b)  & 

c)  & 

d)  & 

Answers

a) Constitutional isomers

b) Same compound

c) Constitutional isomers

d) Different compounds

> See LibreText 7.4 Alkanes and Alkane Isomers

Strategy Map

	Step	Hint
	1. Identify the atoms that make up each compound (their chemical formula).	
ify t…	2. Identify the connectivity of each compound.	Recall the different ways to draw chemical structures and what line drawings of organic compounds represent (see LibreText section 7.2):
ify t…	3. Compare the two compounds and identify what is the same/different between the two.	

Solution

a) Both structures have the same formula but have different connectivity which makes them **constitutional isomers**.

b) Both structures have the same formula and the same connectivity which makes them the **same compound**.

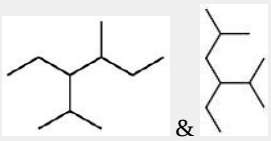
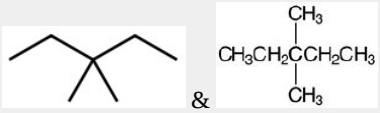
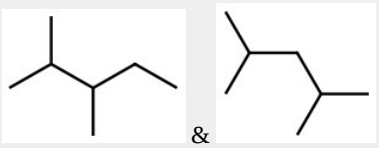
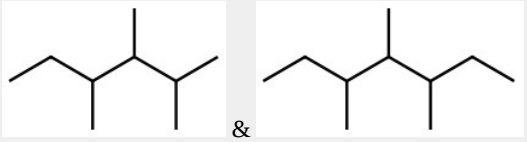
c) Both structures have the same formula but have different connectivity which makes them **constitutional isomers**.

d) The structures have different formulas which makes them **different compounds**.

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory type problem where you must identify structural isomers based on their chemical formula and connectivity.	See LibreText 7.4 Alkanes and Alkane Isomers
tion	Question: Indicate whether each of the following sets are constitutional isomers, the same compound, or different compounds.	
tion	Recall the different ways we can draw chemical structures.	See LibreText 7.2 Drawing Chemical Structures
tion	Recall what a constitutional isomer is.	Constitutional isomers are two or more molecules that have the same chemical formula (are composed of the same atoms), but have different connectivity. For example, the methane group is attached to carbon four on one compound and carbon six on another, but both have the same chemical compounds.
tion	What will make the two compounds different compounds?	The two compounds are considered different compounds if they have a different arrangement or connectivity of atoms.

Guided Solution	Hint
Complete Solution: **a)** Count the C and H atoms. We count Count the C and H atoms. We count seven C and sixteen H for each structure. Both structures therefore have the molecular formula. Both structures therefore have the molecular formula $C_{10}H_{22}$. Looking closely you can see the atoms are connected differently, which makes them constitutional isomers. **Answer** constitutional isomers **b)** Count the C and H atoms. We count seven C and sixteen H for each structure. Both structures therefore have the molecular formula C_7H_{16} and they have the same connectivity which makes them the same compound. **Answer** same compound **c)** Count the C and H atoms. We count seven C and sixteen H for each structure. Both structures therefore have the molecular formula of C_7H_{16} but looking closely we can see they have different connectivity which makes them constitutional isomers. **Answer** constitutional isomers **d)** Count the C and H atoms. We count nine C and twenty H for the left structure and count ten C and twenty-two H for the right structure. The structure on the left has a molecular formula of C_9H_{20} and the structure on the right has a formula of $C_{10}H_{22}$ which makes them different compounds. **Answer** different compounds	 Notice that the C chains in the two structures are different.  These are the same just different ways to represent the structure.  Notice the different position of one of the C branches making the connectivity different even though the atoms are the same.  C_9H_{20} $C_{10}H_{22}$
Check your work! Reviewing your work above checking the connectivity and number of atoms will confirm your answers. **Why do these answers make chemical sense?** The formula of compounds is determined by the atoms that compose them. Multiple compounds can have the same chemical makeup and formula, but their connectivity may be different. If this is the case, they will be constitutional isomers as in (a) and (c) above. This differentiation is made because they will often have different chemical properties. If the number of atoms and connectivity is the same then they will be the same compound as in (b) where you are comparing a line structure drawing to a condensed structural formula. If the number of atoms are different as in (d) the molecular formula is different and they are different compounds.	See LibreText 7.4 Alkanes and Alkane Isomers

(question source from page titled 3.2: Alkanes and Alkane Isomers https://chem.libretexts.org/Bookshelves/Organic_Chemistry/Organic_Chemistry_(Morsch_et_al.)/03%3A_Organic_Compounds-_Alkanes_and_Their_Stereochemistry/3.02%3A_Alkanes_and_Alkane_Isomers shared under a CC BY 4.0 and was authored, remixed, and/or curated by Steven Farmer, Dietmar Kennepohl, Zachary Sharrett, Layne Morsch, Krista Cunningham, & Tim Soderberg.)

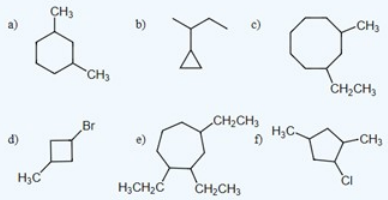
7.3: Question 7.4.6 PASS - identifying constitutional isomers is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.4: Question 7.5.2.1 PASS - naming cycloalkane structures

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 7.5.2.1

Give the IUPAC names for the following cycloalkane structures.



Answer

a) 1,3-Dimethylcyclohexane

b) 2-Cyclopropylbutane

c) 1-Ethyl-3-methylcyclooctane

d) 1-Bromo-3-methylcyclobutane

e) 1,2,4-Triethylcycloheptane

f) 1-Chloro-2,4-dimethylcyclopentane

> See LibreText 7.5.2 Naming Cycloalkanes

Strategy Map

Step	Hint
1. Identify the longest chain (making up the ring or attached to it).	Recall LibreText 7.5.2 Naming Cycloalkanes (see section IUPAC Rules for Nomenclature)
ify t… 2. Count the number of carbons in the ring and along the carbon chains.	
ify t… 3. Number the connecting chains/rings based on alphabetical order.	Ensure that you number each connecting chain going along the path that allows for the lowest numbers (this can be either left or right).

Solution (parts b, c, d and f)

b) The longest carbon chain is in the form of a four-carbon long chain.

Attached to the second carbon of the chain is a three-carbon long ring.

Answer 2-Cyclopropylbutane

c) The longest chain is a eight-carbon long ring.

Attached to the first carbon on the ring is a 2-carbon long chain.

Attached to the third carbon in the ring is a one-carbon long chain.

Answer 1-Ethyl-3-methylcyclooctane

d) The longest chain is a four-carbon long ring.

Attached to the first carbon on the ring is a bromine.

Attached to the third carbon on the ring is a one-carbon long chain.

Answer 1-Bromo-3-methylcyclobutane

f) The longest chain is a five-carbon long ring.

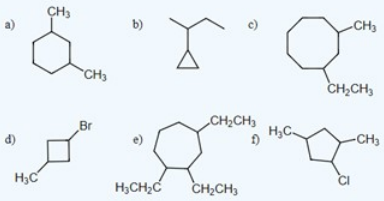
Attached to the first carbon on the ring is a chlorine atom.

Two one-carbon long chains are attached on carbons two and four of the ring.

Answer 1-Chloro-2,4-dimethylcyclopentane

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory type problem where you will need to be able to name different types of cycloalkanes.	See LibreText 7.5.2 Naming Cycloalkanes
Question: Give the IUPAC names for the following cycloalkane structures. **(parts b, c, d and f)** 	

How to name cycloalkane structures:
1. Find which chain is the longest. This can either be inside of the ring or attached to the outside, but you cannot count inside to outside or vice versa.
  ○ This part will go at the end of the name.
2. Identify your connecting chain(s)
  ○ These will be named and numbered based on their order in the alphabet.
  ○ If you have two of the same chains attached to separate carbons on the main chain, they will be listed in order of number (ex, -1, 3-methyl)
	Recall the names for your carbon chains:
	• 1 = methane
	• 2 = ethane
	• 3 = propane
	• 4 = butane
	• 5 = pentane
	• 6 = hexane
	• 8 = octane
	• 9 = nonane
	• 10 = decane
	• If it is a ring, just put “cyclo” in front.

	Guided Solution	Hint
tion	**Complete Solution:** **b)** The longest carbon chain is in the form of a four-carbon long chain. (butane) Attached to the second carbon of the chain is a three-carbon long ring. (2-cyclopropyl) Answer **2-Cyclopropylbutane** **c)** The longest chain is an eight-carbon long ring. (cyclooctane) Attached to the first carbon on the ring is a 2-carbon long chain. (1-ethyl) Attached to the third carbon in the ring is a one-carbon long chain. (3-methyl) Answer **1-Ethyl-3-methylcyclooctane** **d)** The longest chain is a four-carbon long ring. (cyclobutene) Attached to the first carbon on the ring is a bromine. (1-Bromo) Attached to the third carbon on the ring is a one-carbon long chain. (3-methyl) Answer **1-Bromo-3-methylcyclobutane** **f)** The longest chain is a five-carbon long ring. (cyclopentane) Attached to the first carbon on the ring is a chlorine atom. (1-Chloro) Two one-carbon long chains are attached on carbons two and four of the ring. (-2,4-methyl) Answer **1-Chloro-2,4-dimethylcyclopentane**	
tion	**Check your work!** To check if your answer, look back to ensure you followed all of the rules and guidelines. If you accurately named the compound, you should be able to redraw it using your name. The compound you draw should match the original molecule. If it does not, you know you made an error in your naming process. **Why do these answers make chemical sense?** To answer this question you have followed the IUPAC rules for naming cycloalkanes. Key is to make sure your longest chain is the parent part of the name. The longest chain may be a straight chain alkane with the cylcoalkane a substituent. If you focus on identifying the longest chain first and number your carbons to end up with the lowest possible substituent numbers then you have followed the correct approach.	See LibreText 7.5.2 Naming Cycloalkanes (see section IUPAC Rules for Nomenclature)

(question source from page titled 4.1: Naming Cycloalkanes Exercise 4.1.1 https://chem.libretexts.org/Bookshelves/Organic_Chemistry/Organic_Chemistry_(Morsch_et_al.)/04%3A_Organic_Compounds-_Cycloalkanes_and_their_Stereochemistry/4.01%3A_Naming_Cycloalkanes, shared under a CC BY-SA 4.0 license, and was authored, remixed, and/or curated by Jim Clark, Steven Farmer, Dietmar Kennepohl, Zachary Sharrett, Layne Morsch, Krista Cunningham, Tim Soderberg, Pwint Zin, & Kelly Matthews)

7.4: Question 7.5.2.1 PASS - naming cycloalkane structures is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

8: Organic Chemistry II - Stereochemistry

8.1: Question 8.5.2.3 PASS - alkenes, label E or Z

8.2. Question 8.4.E.1 PASS - stereocenter configuration R or S

8.4 Question 8.4.E.2 PASS - enantiomers, diastereomers, identical

---

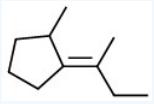
8: Organic Chemistry II - Stereochemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

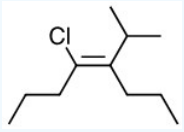
8.1: Question 8.5.2.3 PASS - alkenes, label E or Z

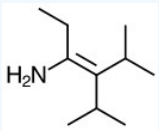
This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise E8.5.2.3

Label the following alkenes as E, Z, or neither

b) 

d) 

e) 

Answers

b) E

d) Z

c) neither

> See LibreText 8.5.2 Sequence Rules - The E,Z Designation

Strategy Map

Step	Hint
1. Look to the left side of the double bond, using the laws for priority, identify if the substituent with the higher priority is on the top or bottom half of the double bond.	Both substituents may have the same priority.
to t… 2. Repeat step one with the right side.	
to t… 3. Identify if the higher priority substituents are both on one side **or** opposite sides of the double bond.	Recall Z has high priority substiuents on same side, and E has high priority substituents on opposite sides (see LibreText 8.5.2 E/Z nomenclature)

Solution

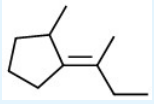
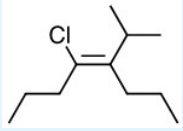
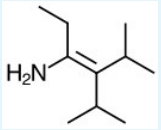
b) E (The substituents with higher priorities are on opposite sides)

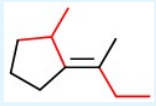
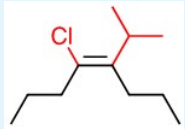
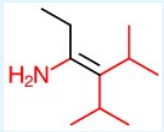
d) Z (The substituents with higher priorities are on the same side)

e) neither (both substituents on the right have the same priority)

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory type problem where you must identify the geometry of alkene structures by labeling them E, Z or neither.	See LibreText 8.5.2 Sequence Rules - The E,Z Designation
tion	Label the following alkenes as E, Z, or neither b)  d)  e) 	
tion	**Recall how to assign substituent priority:** First, determine the two substituents on each double-bonded carbon separately. Rank these substituents based on the atom which is directly attached to the double-bonded carbon. The substituent whose atom has a higher atomic number takes precedence over the substituent whose atom has a lower atomic number. If the atoms are the same, look to the next atom they are bonded to.	Recall rules for assigning priorities LibreText 8.4 Absolute Configuration (Sequence Rules to Assign Priorities to Substituents) See LibreText 8.5.2 Sequence Rules - The E,Z Designation
tion	What does it mean for an alkene to have a type E geometry?	If they have a type E geometry, the highest priority substituents are on opposite sides of the double bond (one is on the top and one is on the bottom).
tion	What does it mean for an alkene to have a type Z geometry?	If they have a type Z geometry, the highest priority substituents are on the same side of the double bond (both on the top or both on the bottom).
tion	What does it mean if they have neither?	If they have neither of these, that means that the substituents on one side (left or right) have the same priority.

	Guided Solution	Hint
tion	**Complete Solution:** part b)  The priority substituents are highlighted in red, and the higher priorities are on opposite sides of the double bond. **answer E** part d)  The priority substituents are highlighted in red, and the higher priorities are on the same side of the double bond. **answer Z** part e)  Both of the substituents showing on the right side of the double bond above have the same priority, so it is neither E or Z. **answer neither**	
tion	**Check your work!** Go back and review your priority assignments to the double bond substituents. Make sure you go atom by atom when looking for priority. **Why does this answer make chemical sense?** The naming rules are used in order to fully describe the molecules when other naming systems do not work. Your answer should make sense if you are following all guidelines to identify priority and geometry.	See LibreText 8.5.2 Sequence Rules - The E,Z Designation

(question source from page titled 7.6: Sequence Rules - The E,Z Designation https://chem.libretexts.org/Bookshelves/Organic_Chemistry/Organic_Chemistry_(Morsch_et_al.)/07%3A_Alkenes-_Structure_and_Reactivity/7.06%3A_Sequence_Rules_-_The_EZ_Designation, shared under a CC BY-SA 4.0 license, authored, remixed, and/or curated by Steven Farmer, Dietmar Kennepohl, Zachary Sharrett, Krista Cunningham, John Roberts, Marjorie C. Caserio & Robert Bruner)

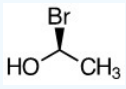
8.1: Question 8.5.2.3 PASS - alkenes, label E or Z is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.2. Question 8.4.E.1 PASS - stereocenter configuration R or S

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 8.4.E.1

For the following compounds, assign R or S configurations for each stereocenter:

a) 

b) 

c) 

Answers

a) S

b) R

c) S

> See LibreText 8.4 Absolute Configuration R-S Sequence Rules

Strategy Map

Step	Hint
1. Assign priorities to substituents and label them one through three.	See LibreText 8.4 Absolute Configuration R-S Sequence Rules
2. Count in the direction of 1 to 3 and identify if this is clockwise or counterclockwise.	Clockwise = R Counterclockwise = S

n pr…

Solution

a) Goes counterclockwise around the stereocenter.

answer S

b) Goes clockwise around the stereocenter.

answer R

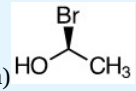
c) Goes counterclockwise around the stereocenter.

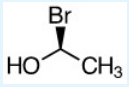
answer S

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint

	Guided Solution	Hint
tion	This is a theory type problem that requires you to assign the correct configuration to a carbon stereocenter (R or S).	See LibreText 8.4 Absolute Configuration R-S Sequence Rules
tion	For the following compounds, assign R or S configurations for each stereocenter: a)  b)  c) 	
tion	The rules to assign priority to substituents: 1. First, examine the atoms directly attached to the stereocenter of the compound. A atom with a higher atomic number takes precedence over a atom with a lower atomic number. Hydrogen is the lowest possible priority atom, because it has the lowest atomic number. 2. If there are two or more substituents which have the same element directly attached to chiral carbon, proceed along the substituent chains until a point of difference is found. Determine which of the chains has the first connection to an atom with the highest priority (the highest atomic number). That chain has the higher priority. 3. For assigning priority, multiple bonds are treated as if each bond of the multiple bond is bonded to a unique atom. For example, an alkene substituent (CH_2=CH-) has higher priority than an ethyl substituent (CH_3CH_2-).	See LibreText 8.4 Absolute Configuration R-S Sequence Rules
tion	R vs S configuration • To identify the configuration of a stereocenter, you must identify if your order of priority goes in the clockwise or counterclockwise direction. • Clockwise = R • Counterclockwise = S	

	Guided Solution	Hint
tion	**Complete Solution:** **a)**  **Bromine** is first Priority (atomic number 35) **Oxygen** is second priority (atomic number 8) **Carbon** is third priority (atomic number 6). Goes counterclockwise around the stereocenter. **answer S** **b)**  **Iodine** is first Priority (atomic number 53) **Fluorine** is second priority (atomic number 9) **Carbon** is third priority (atomic number 6). Goes clockwise around the stereocenter. **answer R** **c)**  **Iodine** is the first priority (atomic number 53), **Oxygen** is the second priority (atomic number 8), **Carbon** is the third priority (atomic number 6). Goes counterclockwise around the stereocenter. **answer S**	
tion	**Check your work!** Check to make sure you are following the guidelines of labeling priority. Ensure you go atom by atom. **Why does this answer make chemical sense?** The "right hand" and "left hand" nomenclature is used to name the enantiomers of a chiral compound, and stereocenters are labeled as (*R*) or (*S*). The naming rules used provide an unambiguous name that indicates the stereochemistry of the molecule.	See LibreText 8.4 Absolute Configuration R-S Sequence Rules

(question source from page titled 6.14: Additional Exercises https://chem.libretexts.org/Courses/Sacramento_City_College/SCC%3A_Chem_420_-_Organic_Chemistry_I/Text/06%3A_Stereochemistry_at_Tetrahedral_Centers/6.14%3A_Additional_Exercises, shared under a CC BY-NC-SA 4.0 license, authored, remixed, and/or curated by Libretexts)

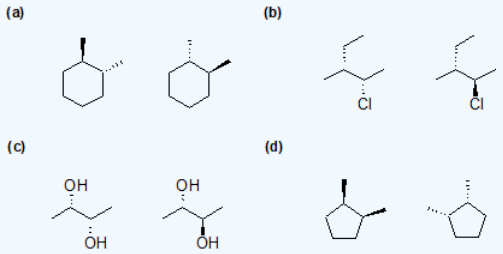
8.2. Question 8.4.E.1 PASS - stereocenter configuration R or S is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.4 Question 8.4.E.2 PASS - enantiomers, diastereomers, identical

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 8.4.E.2

For the following compounds, identify whether they are enantiomers, diastereomers, or the same compound.



Answer

a) Enantiomers

b) Diastereomers

c) Diastereomers

d) Same compound

> See LibreText 8.1 Types of Isomers (Optical Isomers)

Strategy Map

Step	Hint
1. Use your knowledge on compound sketches to “flip” one of the compounds.	When we say “flip” a compound, we are referring to rotating its orientation. This causes atoms that are dashed back to become wedged forward and vice versa. From this, we can see if the two compounds are the same and how they differ.
2. Identify if both compounds have the same shape or are optical isomers.	See LibreText 8.1 Types of Isomers (Optical Isomers)
3. If they are optical isomers identify if they are mirror images of each other.	

Solution

a) Enantiomers

These compounds are mirror image optical isomers.

b) Diastereomers

These compounds are non-mirror image optical isomers.

c) Diastereomers

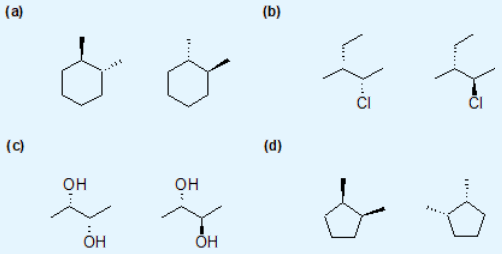
These compounds are non-mirror image optical isomers.

d) Same compound

These compounds have the same chemical formula and connectivity.

Guided Solution

Download Guided Solution as a pdf

	Guided Solution	Hint
tion	This is a theory type problem that requires you to identify if two compounds are optical isomers and what type.	See LibreText 8.1 Types of Isomers (Optical Isomers)
tion	Question - For the following compounds, identify whether they are enantiomers, diastereomers, or the same compound. 	
tion	Recall what these terms mean: • Enantiomer • Diastereomer	**Enantiomers** are non-superimposable mirror images. A common example of a pair of enantiomers is your hands. Your hands are mirror images of one another but no matter how you turn, twist, or rotate your hands, they are not superimposable. **Diastereomers** are non-mirror image optical isomers. Diastereomers have a different arrangement around one or more atoms while some of the atoms have the same arrangement.
tion	Recall how you know if the two compounds are the same.	If two compounds are the same, they will have the same formula and connectivity. When you "flip" the compound it should be superimposable with the other.
tion	Recall how to "flip" the compound.	When we say "flip" a compound, we are referring to rotating its orientation. This causes atoms that are dashed back to become wedged forward and vice versa. From this, we can see if the two compounds are the same and how they differ.

	Guided Solution	Hint
tion	**Complete Solution:** a) These compounds are mirror image optical isomers. These compounds are nonsuperimposable. The two methane groups have opposing orientations (on one compound, the first methane group is wedged and the second is dashed; on the other compound the first methane group is dashed and the second is wedged). **answer Enantiomers** b) These compounds are non-mirror image optical isomers. These are nonsuperimposable compounds, on one compound, the chlorine is dashed back; on the other, the chlorine is wedged forwards. All other connectivity is the same. **answer Diastereomers** c) These compounds are non-mirror image optical isomers. These are nonsuperimposable compounds, on one compound, the hydroxyl group is dashed back; on the other, the hydroxyl group is wedged forwards. All other connectivity is the same. **answer Diastereomers** **d)** These compounds have the same chemical formula and connectivity. They are superimposable. **answer same compound**	
tion	**Check your work!** The easiest way to check optical isomers is to imagine yourself trying to superimpose them (or doing so with a molecular model kit). This will allow you to see which bonds are different. **Why does this answer make chemical sense?** Enantiomers are mirror image optical isomers, and diastereomers are non-mirror image optical isomers. If you flip a compound by rotating its orientation you will end up with superimposable structures if they are the same compound.	Refer to the flowchart Figure 8.11 to check your process.

(question source from page titled 6.14: Additional Exercises https://chem.libretexts.org/Courses/Sacramento_City_College/SCC%3A_Chem_420_-_Organic_Chemistry_I/Text/06%3A_Stereochemistry_at_Tetrahedral_Centers/6.14%3A_Additional_Exercises, shared under a CC BY-NC-SA 4.0 license, authored, remixed, and/or curated by Libretexts)

8.4 Question 8.4.E.2 PASS - enantiomers, diastereomers, identical is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

9: Organic Chemistry III - Conformational Analysis

9.1: Question 9.E.1 PASS - butane Newman projection, bond rotation energy diagram

9.2: Question 9.E.2 PASS - cycloalkane chair, axial vs. equatorial, most stable

---

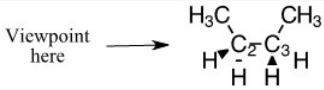
9: Organic Chemistry III - Conformational Analysis is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

9.1: Question 9.E.1 PASS - butane Newman projection, bond rotation energy diagram

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

Exercise 9.E.1

Consider butane's C_2—C_3 bond:



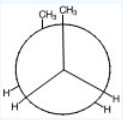
a) Draw the Newman Projection as shown above.

b) Draw the energy profile when the back carbon (C_3) is rotated in 60° clockwise intervals.

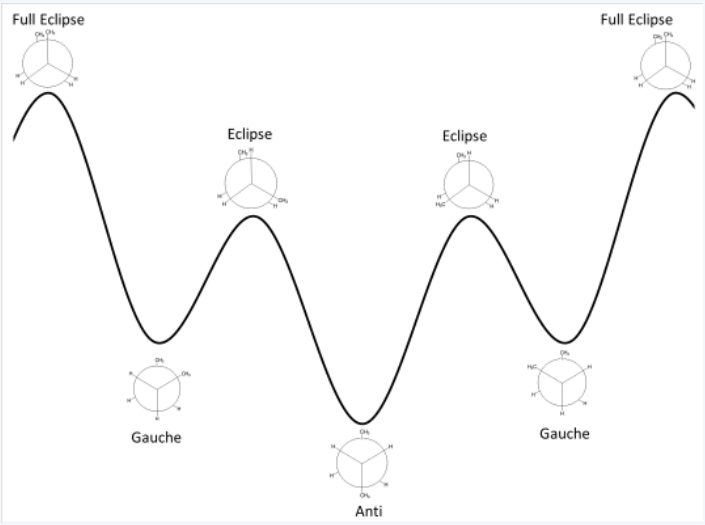
c) Label each conformation.

Answer

a) Newman projection:



b) & c)



See LibreText 9.1 Rotation about single bonds - conformations

Strategy Map

Step	Hint

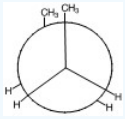
Step	Hint
1. Using the viewpoint given in the question, sketch the Newman Projection of the compound.	

1. Using the viewpoint given in the question, sketch the Newman Projection of the compound.	

2. Sketch each phase as the back carbon (C3) rotates clockwise.	Recall how to draw Newman projections (see LibreText section 9.1)
3. Create a graph depicting the energy change due to the stability of each position.	The closer the atoms are, the more interactions there will be. The further they are, the more stable the compound.
4. Label each position with its corresponding name.	The possible names include Staggered Gauche, Eclipse, Staggered Anti and Full eclipse (not in any order).

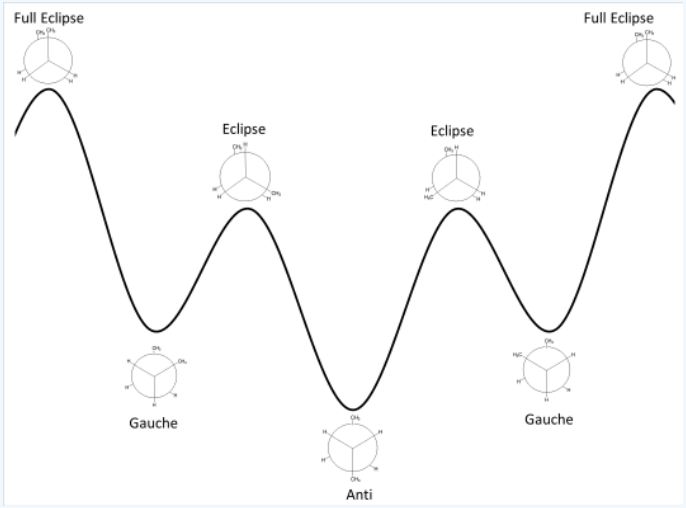
Solution

a)



The exterior circle represents the forward carbon (C_2) and the interior center represents the back carbon (C_3).

b) & c)

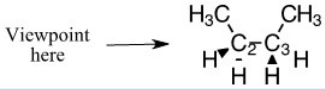
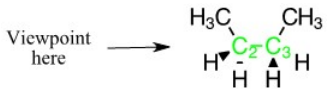
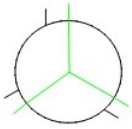
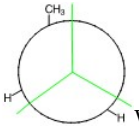
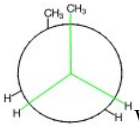
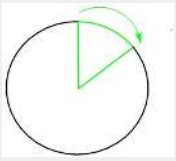
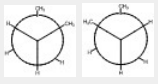
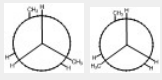


(Least stable) Full eclipse < Eclipse < Staggered Gauche < Staggered Anti (Most stable)

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory type problem that tests your knowledge on the different conformations of a compound and the stability of each conformation.	See LibreText 9.1 Rotation about single bonds - conformations

	Guided Solution	Hint
tion	Question: Consider butane's C_2—C_3 bond:  a) Draw the Newman Projection as shown above. b) Draw the energy profile when the back carbon (C_3) is rotated in 60° clockwise intervals. c) Label each conformation.	
tion	How to sketch a Newman Projection: 1. Identify your viewpoint (which bond are you looking at).	 Looking down the carbon-carbon bond. Carbon two in the front, carbon three in the back.
	2. Draw the skeleton shape of your compound. A line represents the back carbon, and a circle represents the front carbon.	 The back carbon is represented in green, and the front carbon is represented in black.
	3. Draw the groups attached to the front carbon.	 With the substituent groups on the front carbon
	4. Draw the attached groups to the back carbon.	 With the substituent groups on the back carbon.
tion	How much is a 60-degree rotation? There are 360 degrees in a circle, if the bond rotates 60 degrees, it is rotating 1/6 of a circle.	
tion	What impacts the stability of a compounds conformation?	Due to interactions between different atoms, certain positions will have higher stability than others. Certain positions will cause more atoms to be in closer contact and they must overcome more repulsions. This requires energy and therefore causes the compound to be less stable.
tion	Recall the name of each conformation. The possible names include Staggered Gauche, Eclipse, Staggered Anti and Full eclipse (not in any order).	Staggered Gauche  Eclipse  Staggered Anti  Full eclipse 

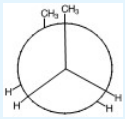
Guided Solution

Hint

tion

Complete Solution:

a) Newman Projection:

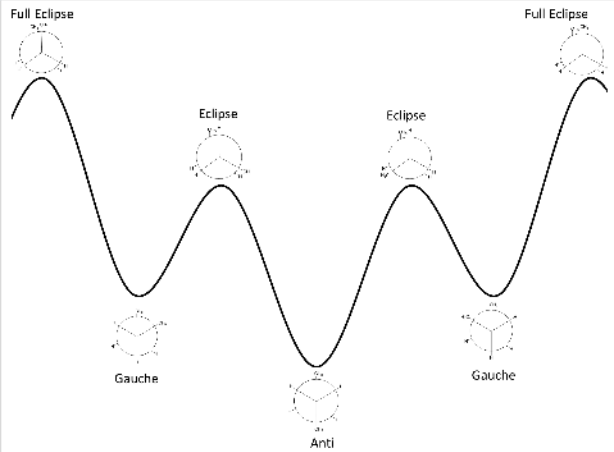


This figure depicts the compound butane down the carbon-carbon bond. The exterior circle represents the forward carbon (C_2) and the interior center represents the back carbon (C_3) The back carbon will rotate at 60° angles.

Step

b) & c)

Energy profile with labeled conformations:



The diagram shows the seven stages of a butane compounds rotation as well as the different levels of stability for each.

(Least stable) Full eclipse < Eclipse < Staggered Gauche < Staggered Anti (Most stable)

The starting phase is Full eclipse, with both of the methane groups side by side. This requires the most energy and therefore is the least stable position.

After the first 60° rotation, the position of the compound is called staggered gauche. The methane group on the back carbon (C3) is now to the left of the methane group on the front carbon (C2). This position is significantly more stable.

After the next 60° rotation, the position of the compound is called eclipse. The methane on the back carbon (C3) is now overlapping with the hydrogen on the front carbon (C2). This position is slightly less stable than the last.

After the next 60° rotation, the position of the compound is called Staggered Anti. The methane group on the back carbon (C3) is now in the opposite position to the methane group on the front carbon (C2). This position is the most stable.

After the next 60° rotation, the position of the compound is called eclipse. The methane on the back carbon (C3) is now

Molecular model of the Newman Projection in Full Eclipse:



Molecular model of the Newman projection in Staggered Gauche:



Molecular model of the Newman projection in Eclipse:



Molecular model of the Newman projection in Staggered Anti:



Molecular model of the Newman projection in eclipse:



Molecular model of the Newman projection in Staggered Gauche:

Guided Solution	Hint
overlapping with the hydrogen on the front carbon (C2). This position is slightly less stable than the last. After the next 60° rotation, the position of the compound is called staggered gauche. The methane group on the back carbon (C3) is now to the right of the methane group on the front carbon (C2). This position is more stable than the previous. After one more 60° rotation, the compound is back to the first position of full eclipse.	
Check your work! Make sure you are working from the point of view in the question. Working with a model is really helpful. **Why does this answer make chemical sense?** Molecules are constantly moving around in 3D space. Single bonds will twist and rotate. Due to interactions between different atoms, certain positions will have higher stability than others. We can analyze the stability of these positions by looking at the atoms in close contact.	See LibreText 9.1 Rotation about single bonds - conformations

(Original question authored by Lindsay Blackstock, shared under a CC BY-NC-SA 4.0 license)

9.1: Question 9.E.1 PASS - butane Newman projection, bond rotation energy diagram is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

9.2: Question 9.E.2 PASS - cycloalkane chair, axial vs. equatorial, most stable

This work by Jensen, Brewer, Blackstock is licensed under CC BY-NC 4.0

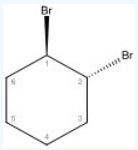
Exercise 9.E.2

Consider: Trans 1,2-dibromocyclohexane

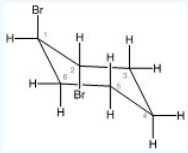
- Draw the simple line structure.
- Draw the two chair confirmations, numbering the carbons and labelling the positions as axial/equatorial.
- Identify which is the more stable structure.

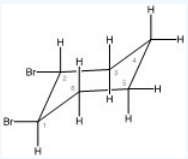
Answer

Simple Line Structure:



Chair conformations:

axial conformation 

equatorial conformation 

The equatorial conformation is the most stable.

> See LibreText 9.4 Cyclohexane - A Strain-Free Cycloalkane

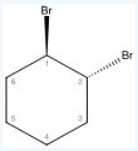
Strategy Map

	Step	Hint
	1. Use the provided molecule name to draw its simple line structure.	Recall drawing chemical structures, line formula (see LibreText section 7.2)
he p…	2. Draw the templates of the two possible conformations.	See LibreText 9.4 Cyclohexane - A Strain-Free Cycloalkane (How to draw Axial and Equatorial bonds)
he p…	3. Number the carbons one through six.	
he p…	4. Label the axial and equatorial positions.	
he p…	5. Convert your simple line structure into your two templates.	

Step	Hint
1. Use the provided molecule name to draw its simple line structure.	Recall drawing chemical structures, line formula (see LibreText section 7.2)
6. Compare the two conformations and identify which would be the most stable.	The most stable conformation is the one that has the least number of interactions between substituent atoms. The larger, more electronegative atoms are favoured in the equatorial positions as they will be impacted by less interactions with surrounding atoms. See LibreText 9.5 Substituted Cyclohexanes

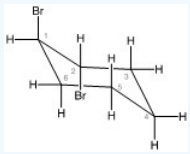
Solution

1. Simple Line Structure:

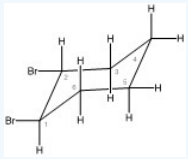


This molecule is trans meaning one bromine is dashed back and one is wedged forward. These bromines are on a six-carbon ring and are attached to carbons one and two.

Chair conformations:

axial 

Both bromines are in axial positions.

equatorial 

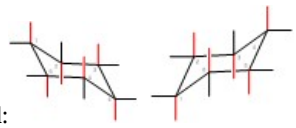
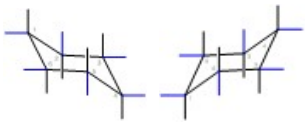
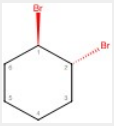
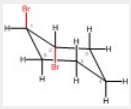
Both bromines are in equatorial positions.

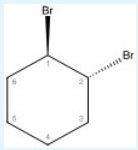
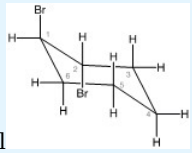
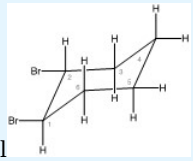
The **equatorial conformation** is the most stable. It has the least interactions between the bromine groups and surrounding atoms.

Guided Solution

Download Guided Solution as a pdf

Guided Solution	Hint
This is a theory type problem that tests your understanding of cycloalkane naming and chair conformations.	See LibreText 9.4 Cyclohexane - A Strain-Free Cycloalkane and see LibreText 9.5 Substituted Cyclohexanes
Consider: Trans 1,2-dibromocyclohexane Draw the simple line structure. Draw the two chair confirmations, numbering the carbons and labelling the positions as axial/equitorial. Identify which is the more stable structure.	

	Guided Solution	Hint
tion	Recall how to sketch a line structure: 1. Identify if the structure is Cis or Trans. 2. Determine if the parent chain is in the form of a chain or ring. (This will be at the end of the name). 3. Identify substituent atoms and which carbons they are connected to.	**Trans** 1,2-dibromocyclohexane This is a cyclohexane, which tells us it is a ring structure.  The numbers tell us the bromines are attached to carbon 1 and carbon 2. **Trans** means one substituent will be dashed back and one will be wedged forward. 
tion	How to draw the chair conformation: 1. Draw two slightly offset parallel lines. 2. Draw another pair of parallel lines from the ends of the first pair. 3. Connect with the third set of parallel lines. 4. To draw its ring-flip conformer, just start the first pair of lines at the opposite angle.	See LibreText 9.4 Cyclohexane - A Strain-Free Cycloalkane (How to Draw the Chair Conformation)
tion	Identify the axial and equatorial positions.	axial positions in red:  equatorial positions in blue: 
tion	Convert your simple line structure to your chair conformation: 1. Match the numbered carbons from your line structure to your template. 2. Recall that substituents wedged forward will be on the higher position and the substituent dashed backwards will be on the lower position.	 to 
tion	Identify the more stable conformation.	The most stable conformation is the one that has the least number of interactions between substituent atoms. The larger, more electronegative atoms are favoured in the equatorial positions as they will be impacted by less interactions with surrounding atoms. See LibreText 9.5 Substituted Cyclohexanes

Guided Solution	Hint
Complete Solution: Simple Line Structure:  This molecule is trans meaning one bromine is dashed back and one is wedged forward. These bromines are on a six-carbon ring and are attached to carbons one and two. Chair conformations:  axial Both bromines are in axial positions. The bromine on carbon one is in the highest position and the bromine on carbon two is in the lowest position.  equatorial Both bromines are in equatorial positions. All axial and equatorial positions switch. The bromine on carbon one is in the highest position and the bromine on carbon two is in the lowest position. is the most stable. It has the least interactions between the bromine groups and surrounding atoms because they are both in equatorial positions. **Answer** The equatorial conformation is the most stable. It has the least interactions between the bromine groups and surrounding atoms because they are both in equatorial positions.	
Check your work! To analyze if your answer is correct, evaluate which conformation has more equatorial positions and given the most space for any larger groups. **Why does this answer make chemical sense?** In the equatorial positions, the larger bromine atoms are the farthest away from all other substituent atoms. The more interactions there are, the more energy it would require and therefore the less stable it would be.	See LibreText 9.5 Substituted Cyclohexanes

(Original question authored by Lindsay Blackstock, shared under a CC BY-NC-SA 4.0 license)

9.2: Question 9.E.2 PASS - cycloalkane chair, axial vs. equatorial, most stable is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

Index

Glossary

Sample Word 1 | Sample Definition 1

Detailed Licensing

Overview

Title: PASS Chemistry Book CHEM 1500

Webpages: 49

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY-NC-SA 4.0: 79.6% (39 pages)
- Undeclared: 20.4% (10 pages)

By Page

- PASS Chemistry Book CHEM 1500 - *CC BY-NC-SA 4.0*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction - *CC BY-NC-SA 4.0*
    - 1.1: About PASS and the Creators - *CC BY-NC-SA 4.0*
    - 1.2: How to use PASS Chem - *CC BY-NC-SA 4.0*
  - 2: Quantum Theory and Electronic Structure - *CC BY-NC-SA 4.0*
    - 2.1: Question 2.E.06 PASS - Energy, wavelength, frequency, and colour of emitted lithium photons - *CC BY-NC-SA 4.0*
    - 2.2: Question 2.E.26 PASS - Bohr Model, quantized energy change - *CC BY-NC-SA 4.0*
    - 2.3: Question 2.E.40 PASS - quantum numbers, orbital characteristics - *CC BY-NC-SA 4.0*
    - 2.4: Question 2.E.58 PASS - electron configuration, orbital diagrams, Aufbau's Principle Hund's Rule - *CC BY-NC-SA 4.0*
  - 3: Periodic Relationships Among the Elements - *CC BY-NC-SA 4.0*
    - 3.1: Question 3.E.05 PASS - ranking first ionization energy - *CC BY-NC-SA 4.0*
    - 3.2: Question 3.E.09 PASS - ranking atomic radius - *CC BY-NC-SA 4.0*
    - 3.3: Question 3.E.18 PASS - describe ionic radii of isoelectronic series - *CC BY-NC-SA 4.0*
  - 4: Chemical Bonding I - Basic Concepts - *CC BY-NC-SA 4.0*
    - 4.1: Question 4.E.20 PASS - which is the most polar bond? - *CC BY-NC-SA 4.0*
    - 4.2: Question 4.E.60 PASS - draw Lewis Structure and calculate formal charge - *CC BY-NC-SA 4.0*
    - 4.3: Question 4.E.65 PASS - determining enthalpy change from bond energies - *CC BY-NC-SA 4.0*
    - PASS Lewis Structure H5P Activity Interactive Video (4.E.60) - *CC BY-NC-SA 4.0*
  - 5: Chemical Bonding II - Molecular Geometry and Hybridization of Atomic Orbitals - *CC BY-NC-SA 4.0*
    - 5.1: Question 5.E.09 PASS - determine electron pair geometry and molecular geometry - *CC BY-NC-SA 4.0*
    - 5.2: Question 5.E.17 PASS - which central atom makes a polar molecule? - *CC BY-NC-SA 4.0*
    - 5.3: Question 5.E.59 PASS - identify central atom hybridization - *CC BY-NC-SA 4.0*
    - 5.4: Question 5.E.78 PASS - using molecular orbital theory and bond order to determine ion charge - *CC BY-NC-SA 4.0*
  - 6: Intermolecular Forces and Liquids and Solids - *CC BY-NC-SA 4.0*
    - 6.1: Question 6.E.07 PASS - define and explain the types of intermolecular forces - *CC BY-NC-SA 4.0*
    - 6.2: Question 6.E.11 PASS - arrange in order of increasing boiling point - *CC BY-NC-SA 4.0*
    - 6.3: Question 6.E.15 PASS - predict melting point - *CC BY-NC-SA 4.0*
  - 7: Organic Chemistry I - Bonding and Structure - *CC BY-NC-SA 4.0*
    - 7.1: Question 7.1.7 PASS - label hybridized atomic orbitals, identify sigma and pi bonds - *CC BY-NC-SA 4.0*
    - 7.2: Question 7.2 PASS - interpret chemical structure line drawing - *CC BY-NC-SA 4.0*
    - 7.3: Question 7.4.6 PASS - identifying constitutional isomers - *CC BY-NC-SA 4.0*
    - 7.4: Question 7.5.2.1 PASS - naming cycloalkane structures - *CC BY-NC-SA 4.0*
  - 8: Organic Chemistry II - Stereochemistry - *CC BY-NC-SA 4.0*

- 8.1: Question 8.5.2.3 PASS - alkenes, label E or Z - *CC BY-NC-SA 4.0*
- 8.2. Question 8.4.E.1 PASS - stereocenter configuration R or S - *CC BY-NC-SA 4.0*
- 8.4 Question 8.4.E.2 PASS - enantiomers, diastereomers, identical - *CC BY-NC-SA 4.0*

9: Organic Chemistry III - Conformational Analysis - *CC BY-NC-SA 4.0*

- 9.1: Question 9.E.1 PASS - butane Newman projection, bond rotation energy diagram - *CC BY-NC-SA 4.0*
- 9.2: Question 9.E.2 PASS - cycloalkane chair, axial vs. equatorial, most stable - *CC BY-NC-SA 4.0*

Back Matter - *Undeclared*

- Index - *Undeclared*
- Glossary - *Undeclared*
- Detailed Licensing - *Undeclared*
- Detailed Licensing - *Undeclared*

Detailed Licensing

Overview

Title: PASS Chemistry Book CHEM 1500

Webpages: 49

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY-NC-SA 4.0: 79.6% (39 pages)
- Undeclared: 20.4% (10 pages)

By Page

- PASS Chemistry Book CHEM 1500 - *CC BY-NC-SA 4.0*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction - *CC BY-NC-SA 4.0*
    - 1.1: About PASS and the Creators - *CC BY-NC-SA 4.0*
    - 1.2: How to use PASS Chem - *CC BY-NC-SA 4.0*
  - 2: Quantum Theory and Electronic Structure - *CC BY-NC-SA 4.0*
    - 2.1: Question 2.E.06 PASS - Energy, wavelength, frequency, and colour of emitted lithium photons - *CC BY-NC-SA 4.0*
    - 2.2: Question 2.E.26 PASS - Bohr Model, quantized energy change - *CC BY-NC-SA 4.0*
    - 2.3: Question 2.E.40 PASS - quantum numbers, orbital characteristics - *CC BY-NC-SA 4.0*
    - 2.4: Question 2.E.58 PASS - electron configuration, orbital diagrams, Aufbau's Principle Hund's Rule - *CC BY-NC-SA 4.0*
  - 3: Periodic Relationships Among the Elements - *CC BY-NC-SA 4.0*
    - 3.1: Question 3.E.05 PASS - ranking first ionization energy - *CC BY-NC-SA 4.0*
    - 3.2: Question 3.E.09 PASS - ranking atomic radius - *CC BY-NC-SA 4.0*
    - 3.3: Question 3.E.18 PASS - describe ionic radii of isoelectronic series - *CC BY-NC-SA 4.0*
  - 4: Chemical Bonding I - Basic Concepts - *CC BY-NC-SA 4.0*
    - 4.1: Question 4.E.20 PASS - which is the most polar bond? - *CC BY-NC-SA 4.0*
    - 4.2: Question 4.E.60 PASS - draw Lewis Structure and calculate formal charge - *CC BY-NC-SA 4.0*
    - 4.3: Question 4.E.65 PASS - determining enthalpy change from bond energies - *CC BY-NC-SA 4.0*
    - PASS Lewis Structure H5P Activity Interactive Video (4.E.60) - *CC BY-NC-SA 4.0*
  - 5: Chemical Bonding II - Molecular Geometry and Hybridization of Atomic Orbitals - *CC BY-NC-SA 4.0*
    - 5.1: Question 5.E.09 PASS - determine electron pair geometry and molecular geometry - *CC BY-NC-SA 4.0*
    - 5.2: Question 5.E.17 PASS - which central atom makes a polar molecule? - *CC BY-NC-SA 4.0*
    - 5.3: Question 5.E.59 PASS - identify central atom hybridization - *CC BY-NC-SA 4.0*
    - 5.4: Question 5.E.78 PASS - using molecular orbital theory and bond order to determine ion charge - *CC BY-NC-SA 4.0*
  - 6: Intermolecular Forces and Liquids and Solids - *CC BY-NC-SA 4.0*
    - 6.1: Question 6.E.07 PASS - define and explain the types of intermolecular forces - *CC BY-NC-SA 4.0*
    - 6.2: Question 6.E.11 PASS - arrange in order of increasing boiling point - *CC BY-NC-SA 4.0*
    - 6.3: Question 6.E.15 PASS - predict melting point - *CC BY-NC-SA 4.0*
  - 7: Organic Chemistry I - Bonding and Structure - *CC BY-NC-SA 4.0*
    - 7.1: Question 7.1.7 PASS - label hybridized atomic orbitals, identify sigma and pi bonds - *CC BY-NC-SA 4.0*
    - 7.2: Question 7.2 PASS - interpret chemical structure line drawing - *CC BY-NC-SA 4.0*
    - 7.3: Question 7.4.6 PASS - identifying constitutional isomers - *CC BY-NC-SA 4.0*
    - 7.4: Question 7.5.2.1 PASS - naming cycloalkane structures - *CC BY-NC-SA 4.0*
  - 8: Organic Chemistry II - Stereochemistry - *CC BY-NC-SA 4.0*

- 8.1: Question 8.5.2.3 PASS - alkenes, label E or Z - *CC BY-NC-SA 4.0*
- 8.2. Question 8.4.E.1 PASS - stereocenter configuration R or S - *CC BY-NC-SA 4.0*
- 8.4 Question 8.4.E.2 PASS - enantiomers, diastereomers, identical - *CC BY-NC-SA 4.0*

9: Organic Chemistry III - Conformational Analysis - *CC BY-NC-SA 4.0*

- 9.1: Question 9.E.1 PASS - butane Newman projection, bond rotation energy diagram - *CC BY-NC-SA 4.0*
- 9.2: Question 9.E.2 PASS - cycloalkane chair, axial vs. equatorial, most stable - *CC BY-NC-SA 4.0*

Back Matter - *Undeclared*

- Index - *Undeclared*
- Glossary - *Undeclared*
- Detailed Licensing - *Undeclared*
- Detailed Licensing - *Undeclared*