Chapter 3 Review Problems

INSTRUCTIONS:

You ***do not*** need to write the **question**, ONLY WRITE THE PROBLEM NUMBER and ANSWERS/SOLUTIONS.

- For problems that involve calculations, you must ***show your work*** to get full credit.
- For multiple choice questions, you can simply write the letter (a, b, c, or d) of the correct response.
- Use the *navigation buttons* at the bottom of the pages to get hints, check your answers, move to the next problem, or go back to previous pages.

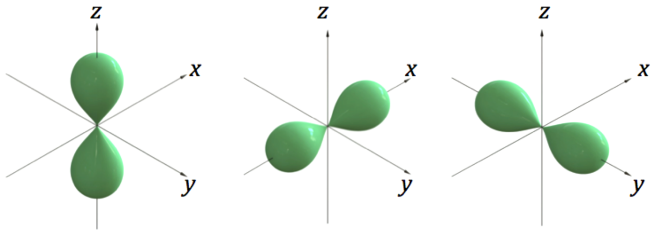
Chapter Review Problems are **due** at the ***end of class period*** on the dates shown in the CHEM 108 Schedule.

- Late submissions will not be accepted unless the student can prove to the instructor that something outside of their control prevented them from turning in the problem set on the due date (see the course syllabus for more details).

Go to first question

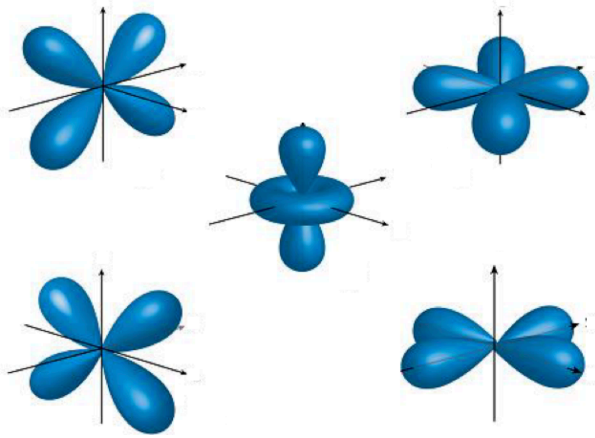
3.1) Classify the atomic orbitals as either **s**, **p**, or **d**.

a)



b)

c)



Go back

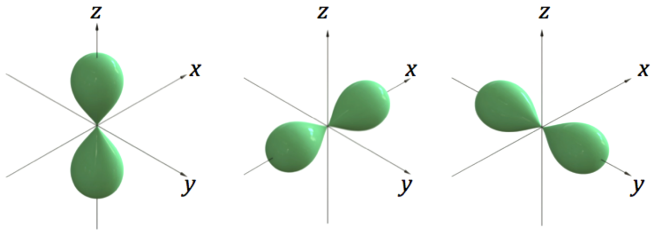
Click here for a **hint**

Click here to **check your answer**

Go to next question

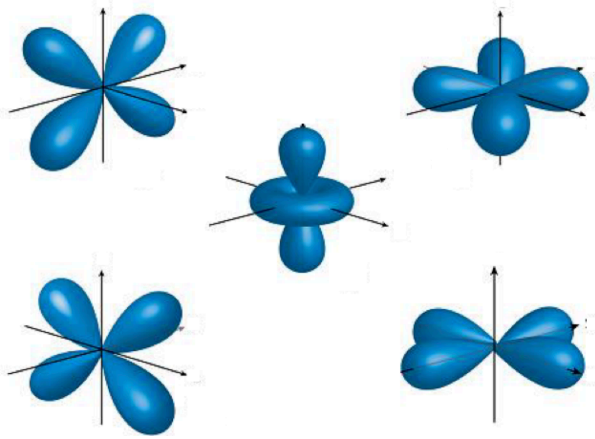
3.1) Classify the atomic orbitals as either **s**, **p**, or **d**.

a)



b)

c)



HINT: Atomic orbitals are classified as s, p, or d based on their shape.

For more help: See chapter 3 part 2 video or chapter 3 section 2 in the textbook.

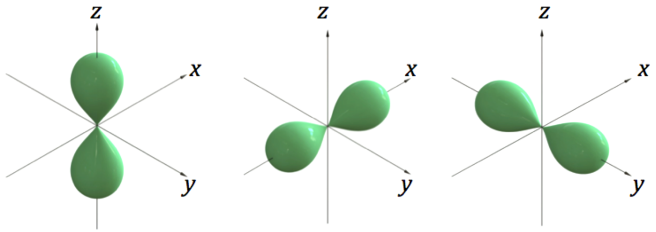
Go back

Click here to **check your answer**

Go to next question

3.1) Classify the atomic orbitals as either s, p, or d.

a)

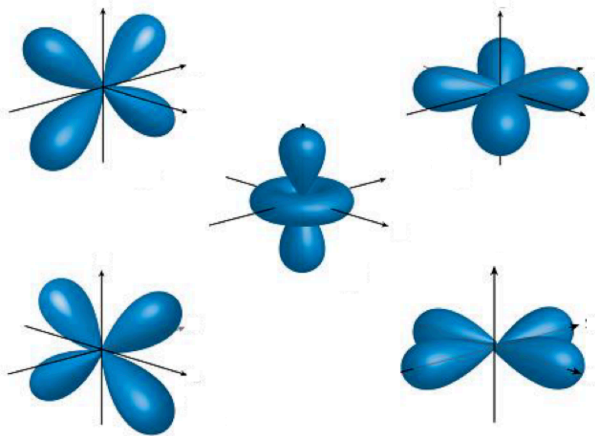


p orbitals

b)

s orbital

c)



d orbitals

EXPLANATION:

Atomic orbitals are classified as s, p, or d based on their shape.

For more details: See chapter 3 part 2 video or chapter 3 section 2 in the textbook.

Go back

Go to next question

3.2) Draw an “energy level diagram” for a nitrogen atom.

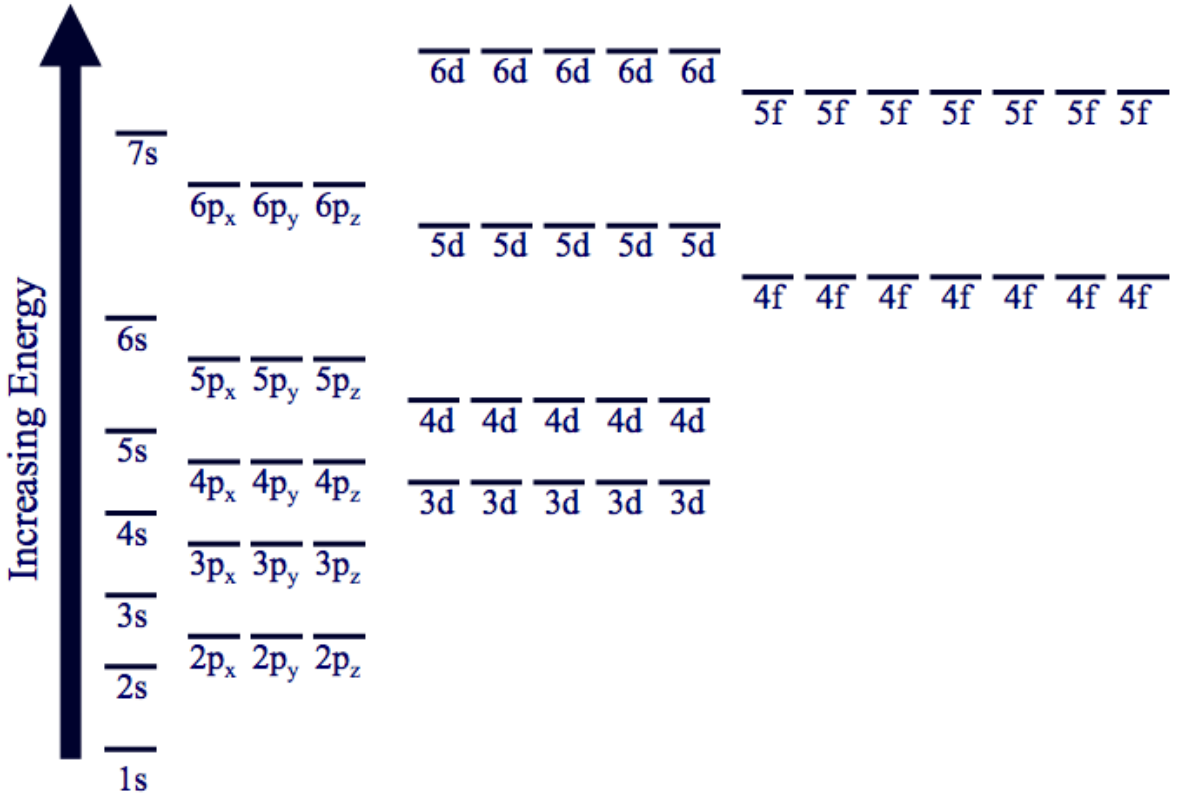
Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.2) Draw an "energy level diagram" for a nitrogen atom.



HINT: A nitrogen atom 7 electrons. An empty (no electrons) energy level diagram for multi-electron atoms is shown on the left. Electrons are arranged (configured) into the orbitals in the way that results in the lowest possible energy. Nature does this by obeying the following principles:

1) The Aufbau Principle:" "Aufbau" (German) means *build-up* or *construct*. The aufbau principle states that an electron occupies the ***lowest energy orbital that can receive it***.

2) The Pauli Exclusion Principle: An orbital can hold a maximum of **two** electrons. *When two electrons occupy the same orbital, one electron has spin "***up***" the other has spin "***down***."* Having two electrons in the same orbital with opposite spin states is lower in energy than when both spins are the same.

3) Hund's Rules : (i) When electrons are configured into orbitals that all have the same energy, for example the $2\mathbf{p}_x$, $2\mathbf{p}_y$, and $2\mathbf{p}_z$, a ***single electron*** is placed into **each** of the equal-energy orbitals before a second electron is added to an occupied orbital. **(ii)** When electrons are configured into a *set of orbitals* that all have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all "up").

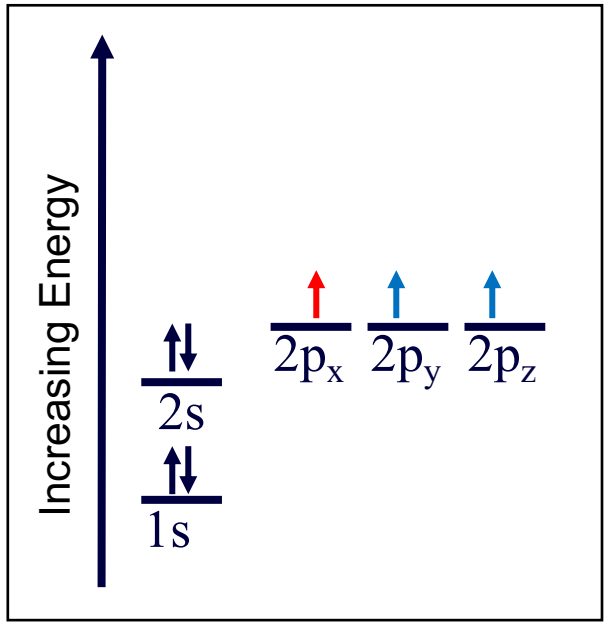
Go back

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Click here to **check your answer**

Go to next question

3.2) Draw an "energy level diagram" for a nitrogen atom.



EXPLANATION: A nitrogen atom has 7 protons and therefore 7 electrons. We use the *multi-electron* energy level diagram.

- *The **first** and **second electrons*** occupy the 1**s** orbital according to the ***Aufbau Principle***.
  - Note that the nitrogen energy level diagram in does not include many high energy orbitals (**n** >2) because nitrogen *only has seven electrons*.
- Because the 1**s** orbital is now full (orbitals hold a maximum of two electrons), *the **third** and **fourth electrons*** occupy the 2**s** orbital.
- The ***fifth electron*** goes into any one of the three 2**p** orbitals. The three 2**p** orbitals all are equivalent in energy. We will arbitrarily choose to put the ***fifth electron*** into the 2**p**$_x$ orbital (represented by the red arrow).

- The ***sixth*** and ***seventh electrons*** (represented by the blue arrows) each go into *unoccupied* 2**p** orbitals. Recall **Hund's Rule (i)**: when electrons are configured into orbitals that all have the same energy, a ***single electron*** is placed into **each** of the equal-energy orbitals before the electrons are added to an occupied orbital.

- How did we know to put the spin state of the sixth and seventh electrons in the "up" state? **Hund's Rule (ii)** explains that when electrons are configured into a *set of orbitals* that have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all "up").

For more details: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Go to next question

3.3) Draw an “energy level diagram” for a calcium (Ca) atom.

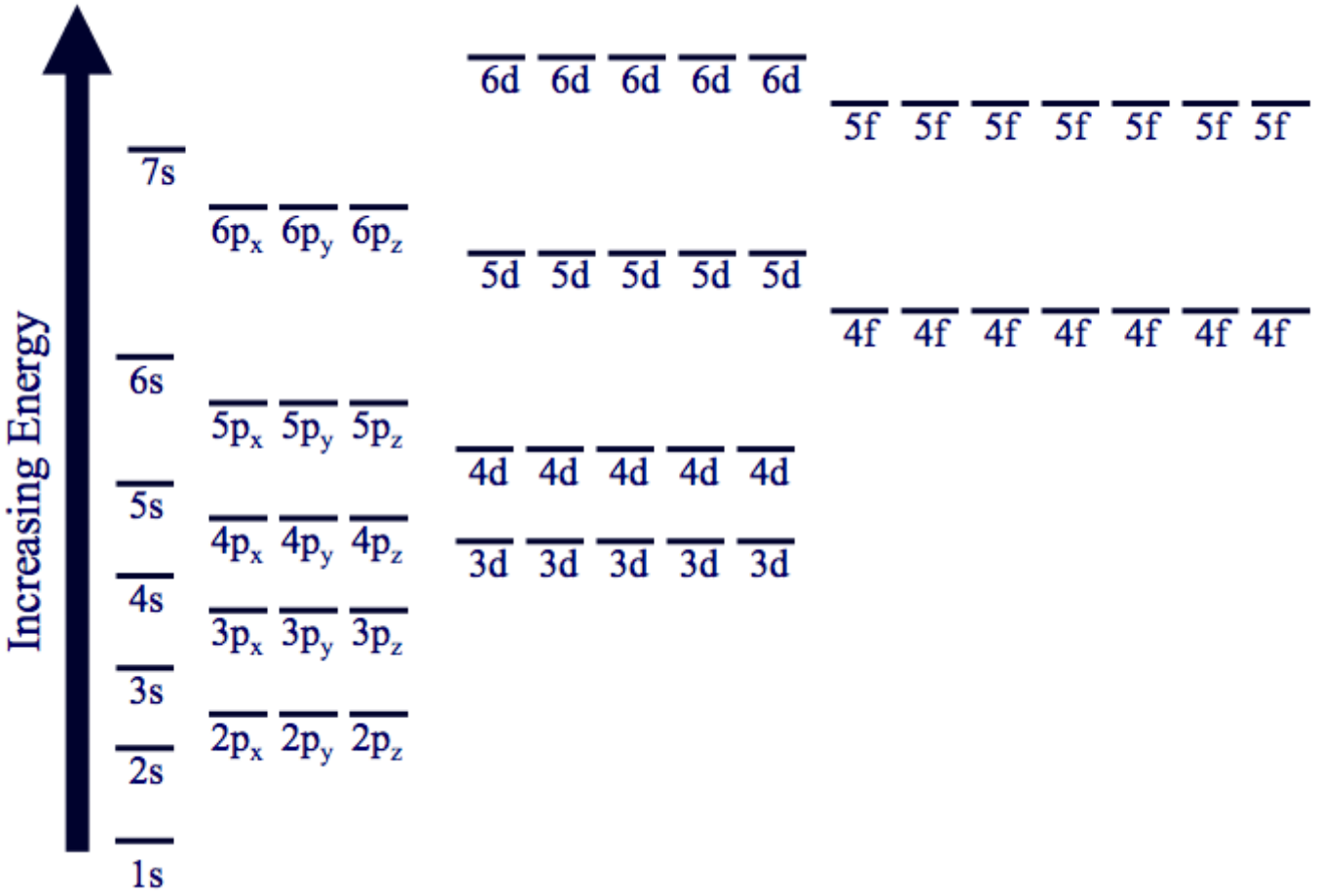
Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.3) Draw an "energy level diagram" for a calcium (Ca) atom.



HINT: Use the periodic table to determine the number of electrons that are present in a krypton atom. An empty (no electrons) energy level diagram for multi-electron atoms is shown on the left. Electrons are arranged (configured) into the orbitals in the way that results in the lowest possible energy. Nature does this by obeying the Aufbau Principle, the Pauli Exclusion Principle, and Hund's Rules. See the hint from the previous question (3.2) for more details.

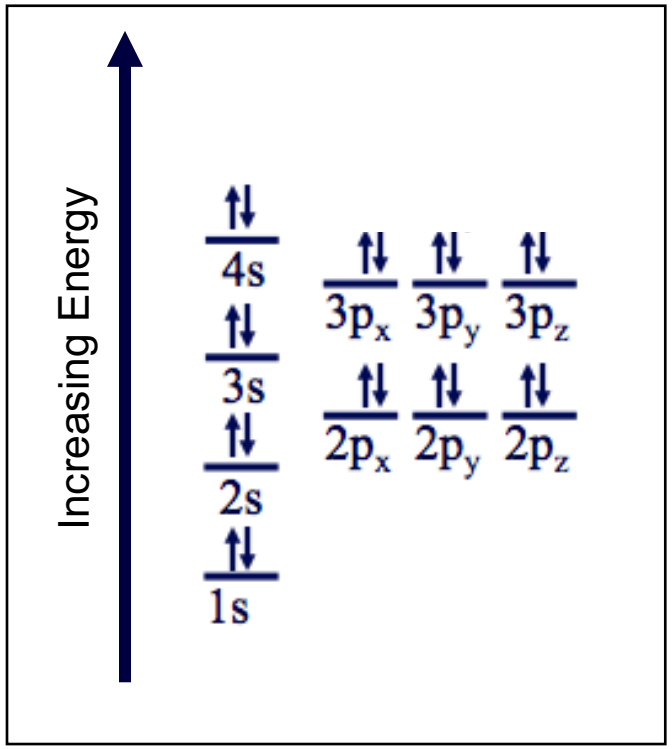
For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.3) Draw an "energy level diagram" for a calcium (Ca) atom.



EXPLANATION: A calcium atom has 20 protons and therefore 20 electrons. Electrons are arranged (configured) into the orbitals in the way that results in the lowest possible energy. Nature does this by obeying the Aufbau Principle, the Pauli Exclusion Principle, and Hund's Rules. See the hint from the previous question (3.2) for more details. Note that the 4s orbital fills before the 3d orbital because the 4s orbital lower in energy than the 3d orbitals (the 3d orbitals are not shown here because they do not contain electrons).

For more details: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Go to next question

3.4) In a situation where an electron goes from the n = 1 orbital (ground state) to an n = 3 orbital, which of the following statements are true?

a) Energy was released by the atom.

b) Energy was absorbed by the atom.

c) There was no change in the atom's energy.

Go back

Click here for a **hint**

Click here to **check your answer**

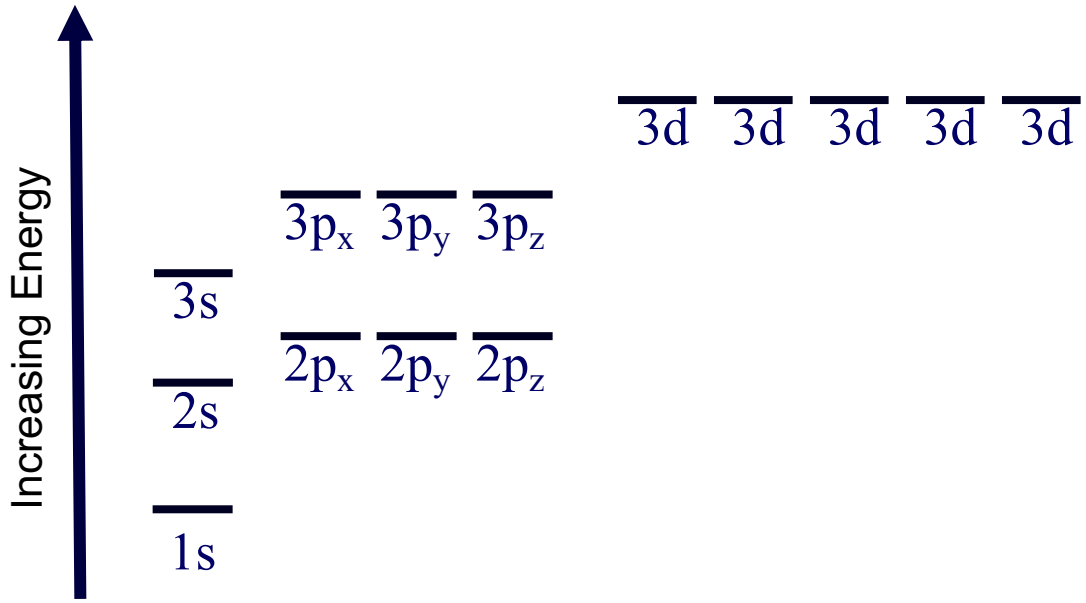
Go to next question

3.4) In a situation where an electron goes from the n = 1 orbital (ground state) to an n = 3 orbital, which of the following statements are true?

a) Energy was released by the atom.

b) Energy was absorbed by the atom.

c) There was no change in the atom's energy.



HINT: Look at the energy level diagram. Base your answer on the relative energy of the n=1 orbital (1s) to any of the n = 3 orbitals (3s, 3p, or 3d).

Go back

Click here to **check your answer**

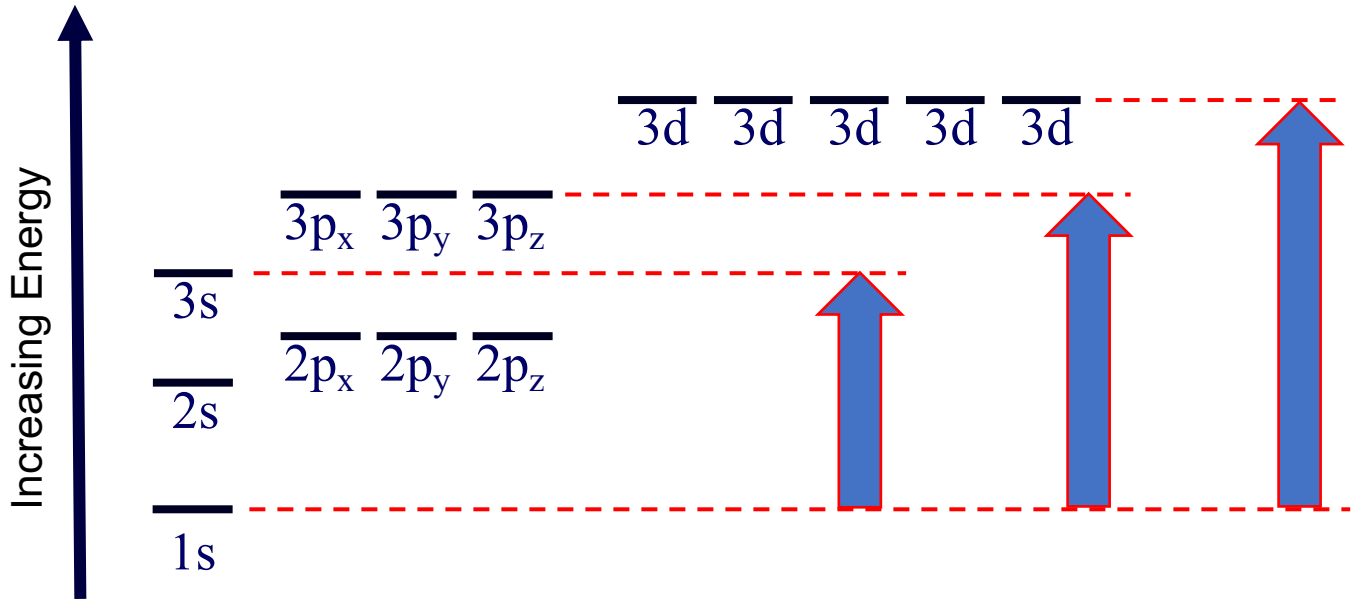
Go to next question

3.4) In a situation where an electron goes from the n = 1 orbital (ground state) to an n = 3 orbital, which of the following statements are true?

a) Energy was released by the atom. **false**

b) Energy was absorbed by the atom. **true**

c) There was no change in the atom's energy. **false**



EXPLANATION:

An atom must **absorb** a quantity of energy indicated by one of the **blue arrows** in order to move from the n=1 orbital to an n = 3 orbital.

Go back

Go to next question

3.5) Label each of these orbital designations as "**possible**" (exists) **or** "**impossible**" (does not exist).

a) 1s

b) 2s

c) 1p

d) 2p

e) 2d

f) 3d

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.5) Label each of these orbital designations as "**possible**" (exists) **or** "**impossible**" (does not exist).

a) 1s

b) 2s

c) 1p

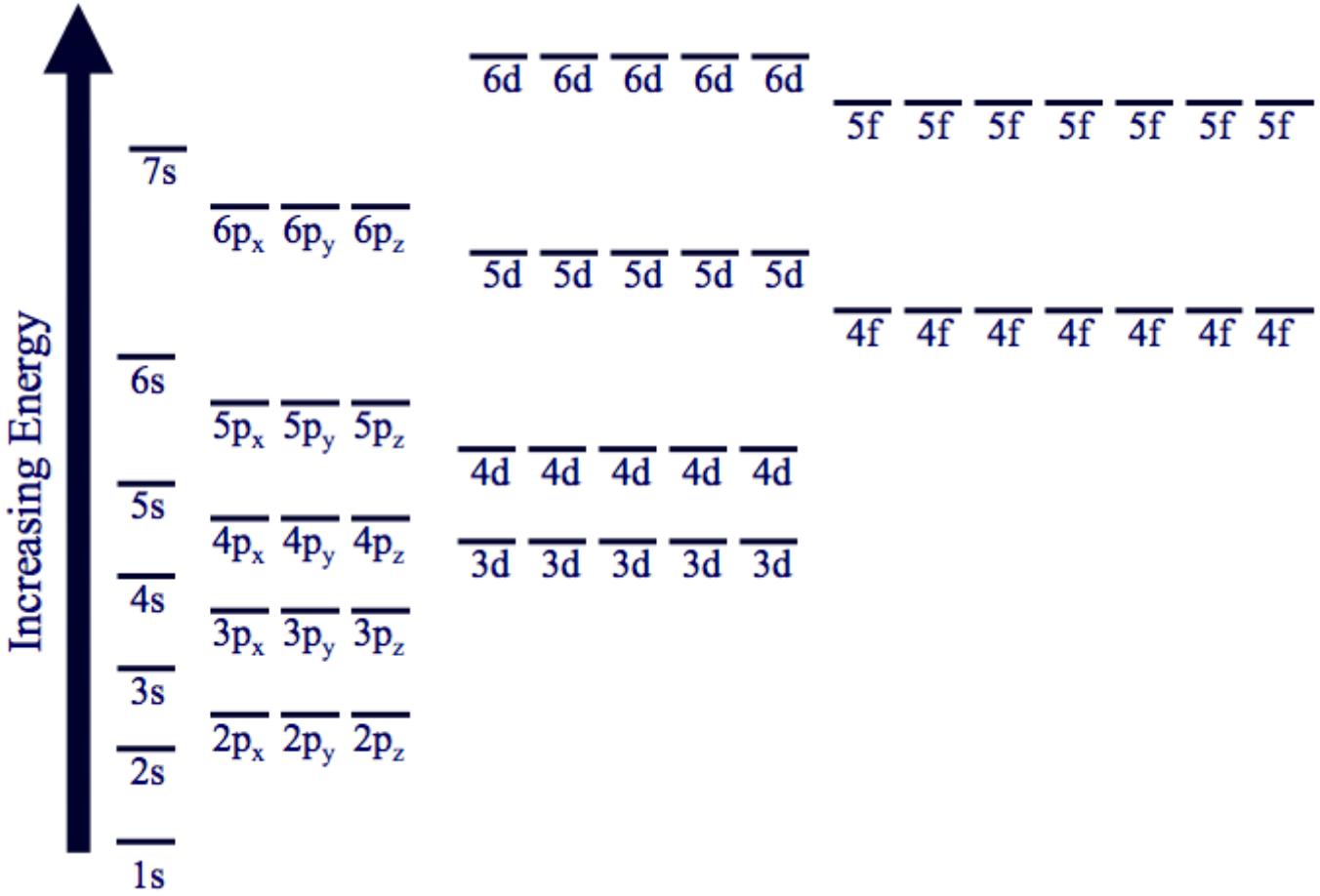
d) 2p

e) 2d

f) 3d

HINT:

Check the energy level diagram to see if these orbitals are present **(possible)** or if they do not exist **(impossible).**



For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.5) Label each of these orbital designations as "**possible**" (exists) **or** "**impossible**" (does not exist).

a) 1s **possible**

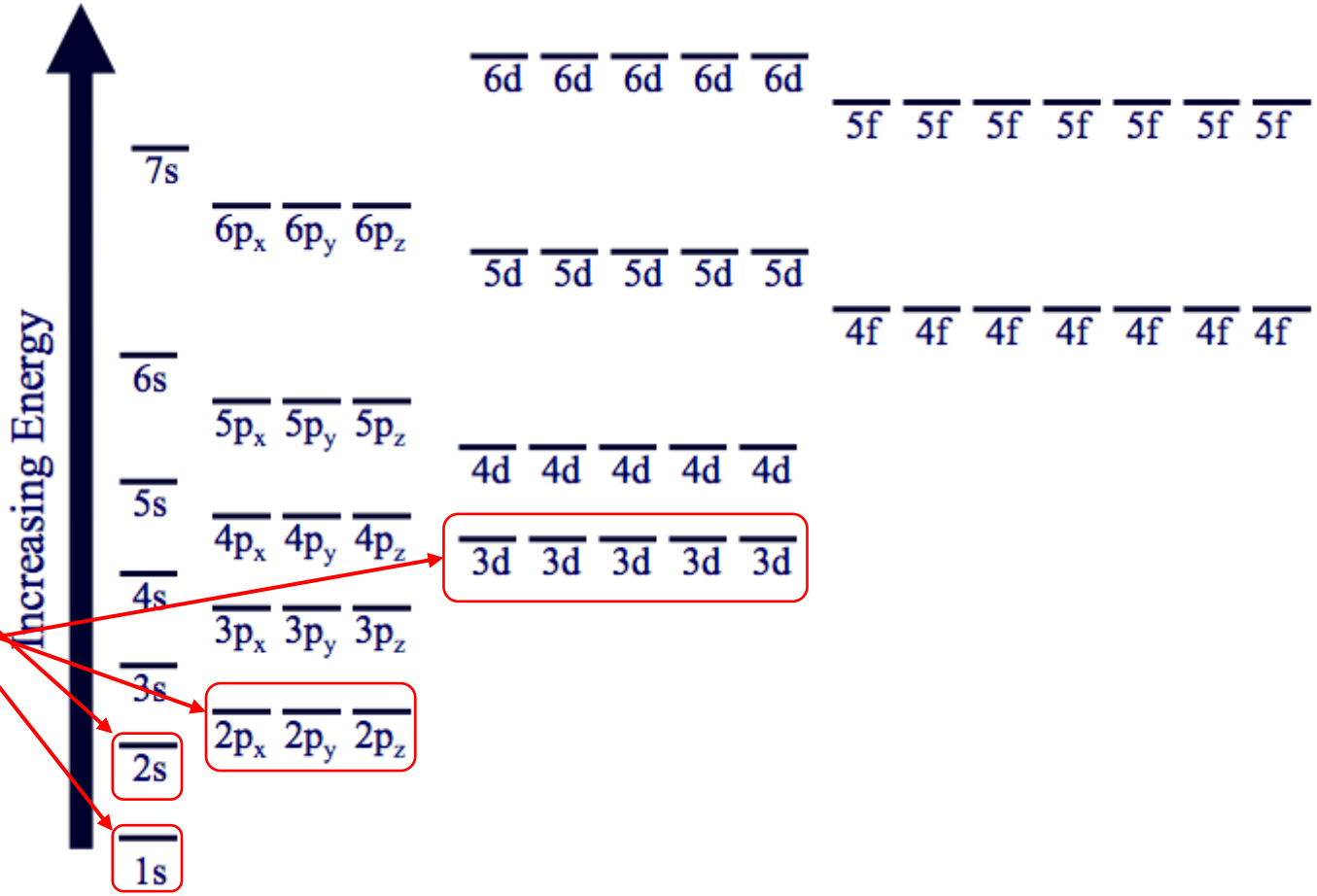
b) 2s **possible**

c) 1p **impossible**
1p orbitals do not exist

d) 2p **possible**

e) 2d **impossible**
2d orbitals do not exist

f) 3d **possible**



For more details: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Go to next question

3.6) Redraw the table below and then enter the number of electrons in each quantum level (**n**) for the following elements. If the quantum level does not contain any electrons, enter a "**0**" (zero).

Element	n = 1	n = 2	n = 3	n = 4
hydrogen (H)				
magnesium (Mg)				
bromine (Br)				
potassium (K)				

Go back

Click here for a **hint**

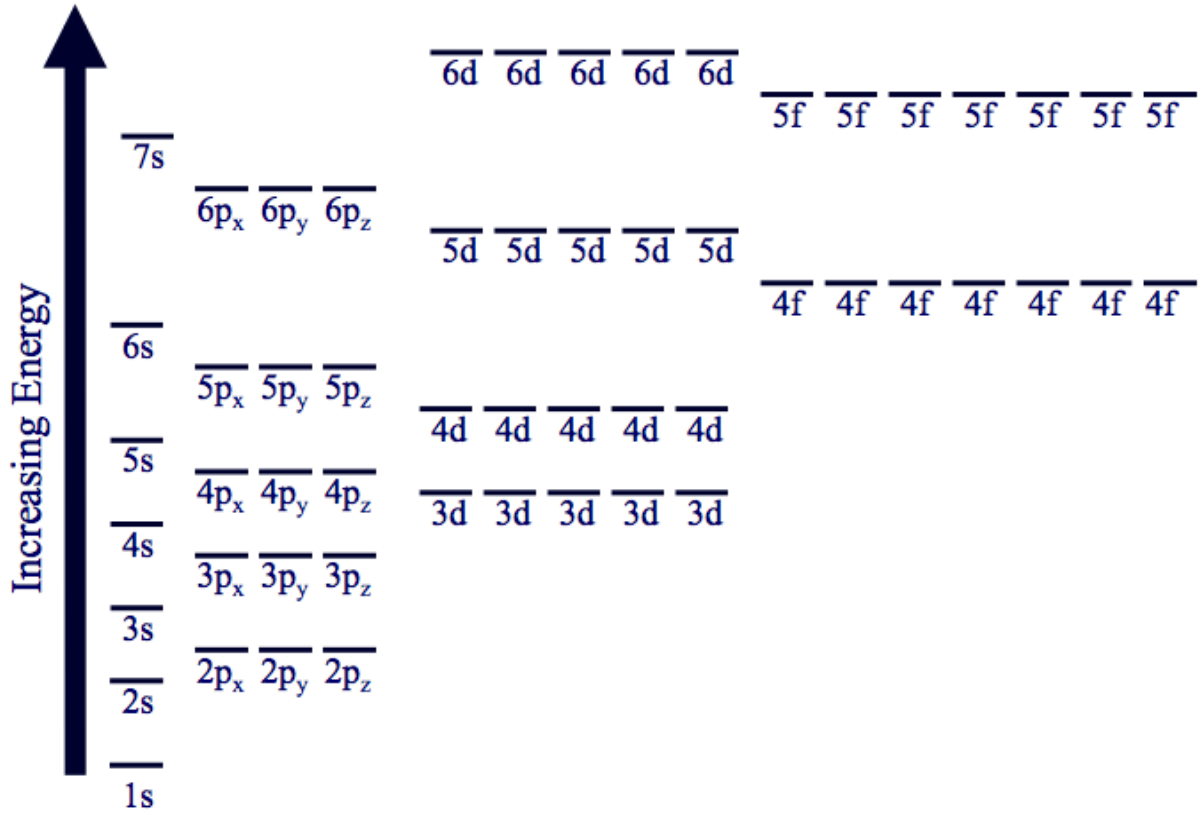
Click here to **check your answer**

Go to next question

3.6) Redraw the table below and then enter the number of electrons in each quantum level (**n**) for the following elements. If the quantum level does not contain any electrons, enter a "**0**" (zero).

HINT: For each element in the table, arrange the electrons into the energy level diagram. Next, count the number of electrons in each quantum level (n).

Element	n = 1	n = 2	n = 3	n = 4
hydrogen (H)				
magnesium (Mg)				
bromine (Br)				
potassium (K)				



For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Click here to **check your answer**

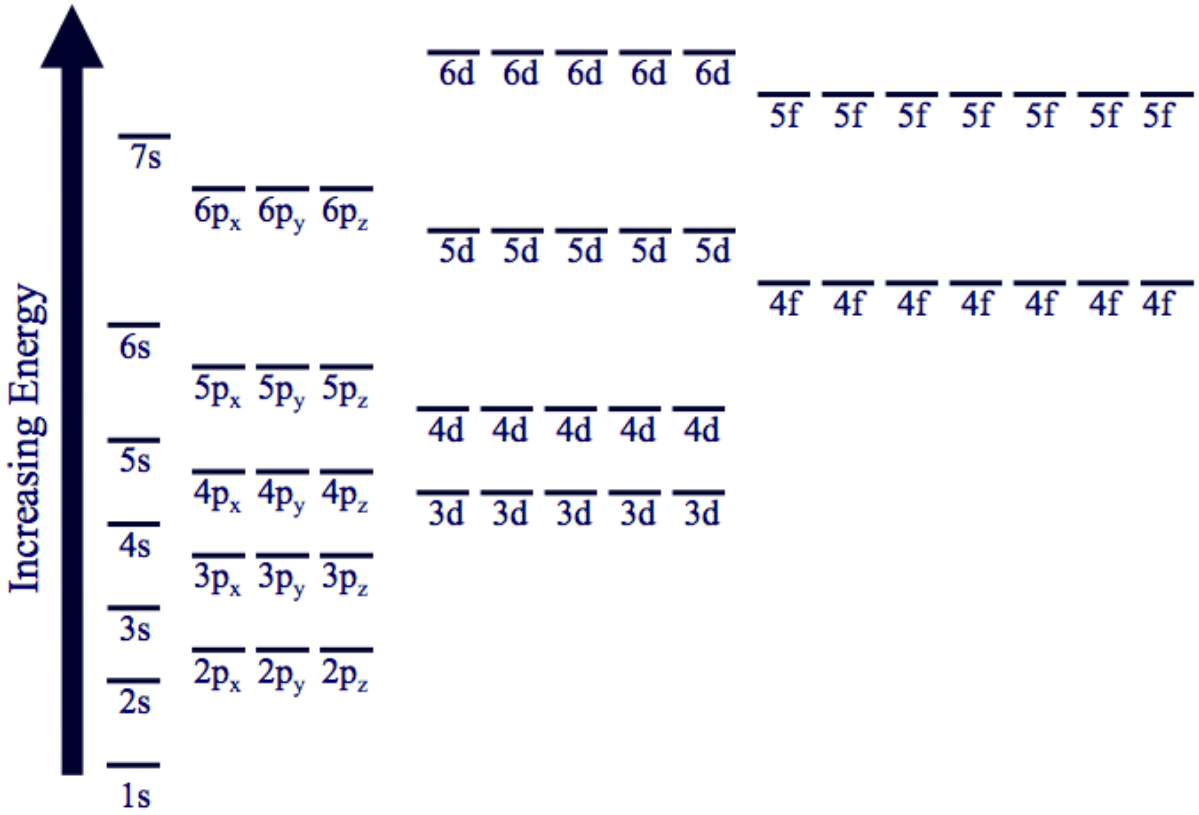
Go to next question

3.6) Redraw the table below and then enter the number of electrons in each quantum level (**n**) for the following elements. If the quantum level does not contain any electrons, enter a "**0**" (zero).

EXPLANATION: For each element in the table, electrons are placed into the energy level diagram. After doing so, you can count the number of electrons in each quantum level (n).

Element	n = 1	n = 2	n = 3	n = 4
hydrogen (H)	**1**	**0**	**0**	**0**
magnesium (Mg)	**2**	**8**	**2**	**0**
bromine (Br)	**2**	**8**	**18**	**7**
potassium (K)	**2**	**8**	**8**	**1**

Note that the 4s orbital is *lower in energy* than the 3d orbitals, therefore the 4s orbital will accept one electron **before** electrons are placed in the 3d orbitals.



Go back

Go to next question

3.7) Determine the number of **valence electrons** in each of the atoms listed below.

a) a hydrogen (H) atom

b) a carbon (C) atom

c) an argon (Ar) atom

d) a sulfur (S) atom

e) a barium (Ba) atom

f) a chlorine (Cl) atom

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.7) Determine the number of **valence electrons** in each of the atoms listed below.

a) a hydrogen (H) atom

b) a carbon (C) atom

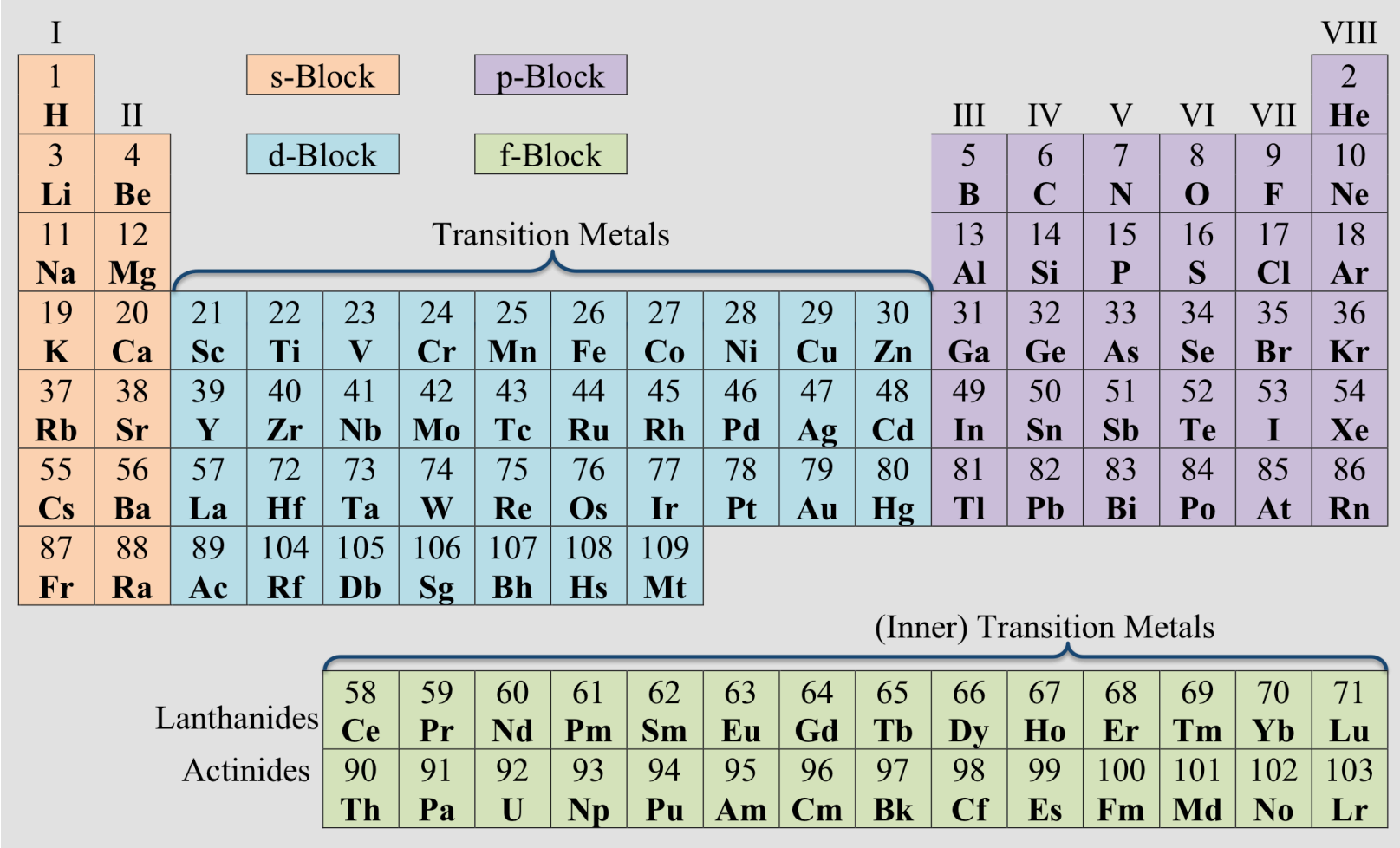
c) an argon (Ar) atom

d) a sulfur (S) atom

e) a barium (Ba) atom

f) a chlorine (Cl) atom

HINT: For *s-block* and *p-block* elements, the number of valence electrons can be determined by the element's *position in the periodic table*.



I	II										III	IV	V	VI	VII	VIII
1 H																2 He
3 Li	4 Be										5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg										13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt									

Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Go back

For more help: See chapter 3 part 4 video or chapter 3 section 3 in the textbook.

Click here to **check your answer**

Go to next question

3.7) Determine the number of **valence electrons** in each of the atoms listed below.

a) a hydrogen (H) atom **1**

b) a carbon (C) atom **4**

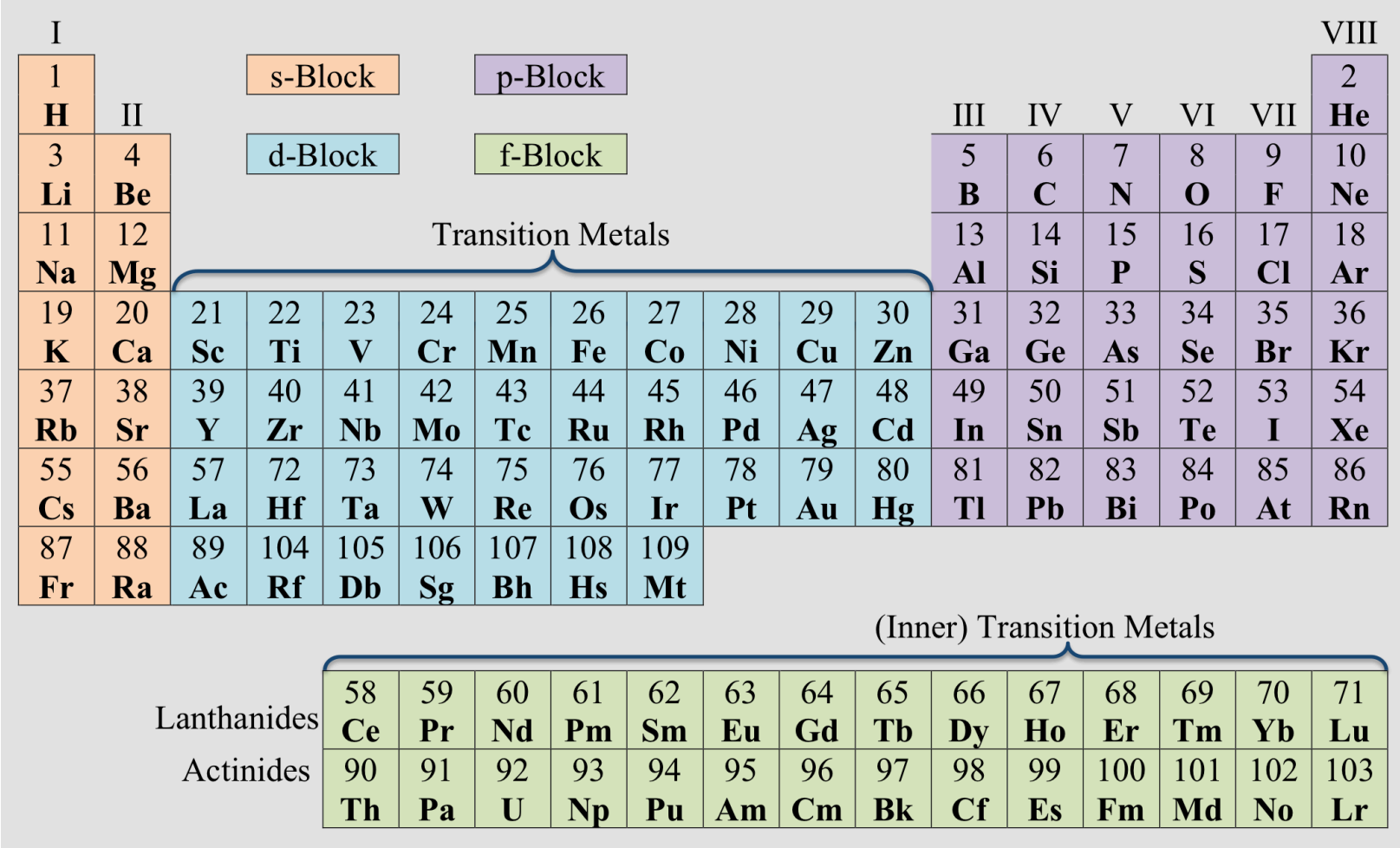
c) an argon (Ar) atom **8**

d) a sulfur (S) atom **6**

e) a barium (Ba) atom **2**

f) a chlorine (Cl) atom **7**

EXPLANATION: For *s-block* and *p-block* elements, the number of valence electrons is the same as the element's *column group number.* The group numbers are shown as roman numerals above the column of elements.



For more details: See chapter 3 part 4 video or chapter 3 section 3 in the textbook.

Go back

Go to next question

3.8) Determine whether the following atoms will **gain** or **lose** electron when forming an ion.

a) an oxygen (O) atom

b) a calcium (Ca) atom

c) a fluorine (F) atom

d) a sodium (Na) atom

e) a nitrogen (N) atom

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.8) Determine whether the following atoms will **gain** or **lose** electron when forming an ion.

a) an oxygen (O) atom

b) a calcium (Ca) atom

c) a fluorine (F) atom

d) a sodium (Na) atom

e) a nitrogen (N) atom

HINT:

Metal atoms ***lose*** electrons to form ***positive ions***.

Nonmetal atoms ***gain*** electrons to form ***negative ions***.

For more help: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.8) Determine whether the following atoms will **gain** or **lose** electron when forming an ion.

a) an oxygen (O) atom ***gain***

b) a calcium (Ca) atom ***lose***

c) a fluorine (F) atom ***gain***

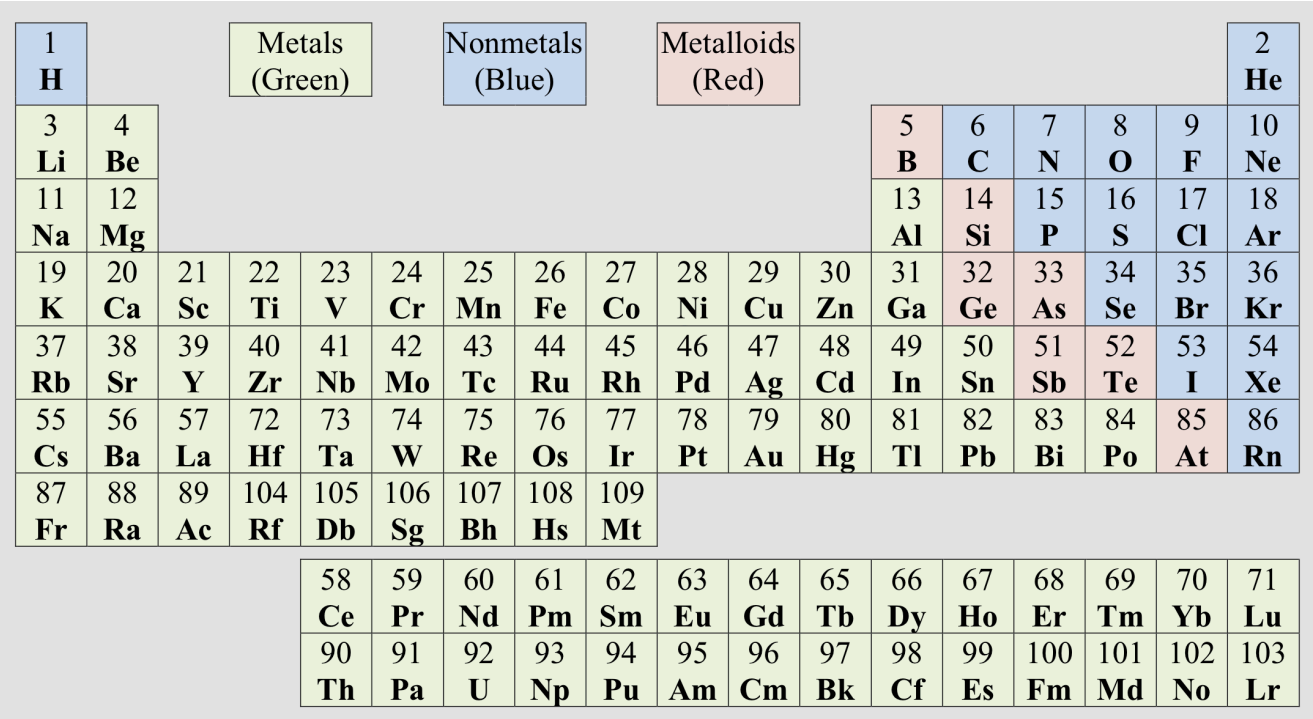
d) a sodium (Na) atom ***lose***

e) a nitrogen (N) atom ***gain***

EXPLANATION:

Metal atoms ***lose*** electrons to form ***positive ions***. If an atom *loses* one or more electrons, it will then have more protons than electrons and have an overall *positive charge*. *Positive ions* are called **cations**.

Nonmetal atoms ***gain*** electrons to form ***negative ions***. If an atom *gains* one or more electrons, it will then have more electrons than protons and have an overall *negative charge*. *Negative ions* are called **anions**.



1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt									
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Go back

For more details: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.

Go to next question

3.9) Predict the charge of the following ions.

a) a sodium ion

b) an oxide ion

c) a calcium ion

d) a fluoride ion

e) a potassium ion

f) a nitride ion

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.9) Predict the charge of the following ions.

a) a sodium ion

b) an oxide ion

c) a calcium ion

d) a fluoride ion

e) a potassium ion

f) a nitride ion

HINT: Ions generally form such that the ***ion*** has an **octet of electrons** in its outermost shell. This tendency will allow us to predict the charge of the ion that is formed for particular elements.

For more help: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.

Charge for s-Block Cations and p-Block Nonmetal Anions

Periodic Group	Number of Valence Electrons of the Element	Number of Electrons Gained or Lost in Ion Formation	Charge of Ion Formed
s-Block Elements			
Group I	1	Lose 1 electron	1+
Group II	2	Lose 2 electrons	2+
p-Block Nonmetal Elements			
Group III	There are no Group III non-metals (only metals and metalloids)		
Group IV	4	Do not form ions, high energy to gain **or** lose 4 electrons!	
Group V	5	Gain 3 electrons	3-
Group VI	6	Gain 2 electrons	2-
Group VII	7	Gain 1 electron	1-
Group VIII	8	Do not form ions, noble gas atoms have filled outer shells.	

Go back

Click here to **check your answer**

Go to next question

3.9) Predict the charge of the following ions.

a) a sodium ion **1+**

b) an oxide ion **2-**

c) a calcium ion **2+**

d) a fluoride ion **1-**

e) a potassium ion **1+**

f) a nitride ion **3-**

Charge for s-Block Cations and p-Block Nonmetal Anions

Periodic Group	Number of Valence Electrons of the Element	Number of Electrons Gained or Lost in Ion Formation	Charge of Ion Formed
s-Block Elements			
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p-Block Nonmetal Elements			
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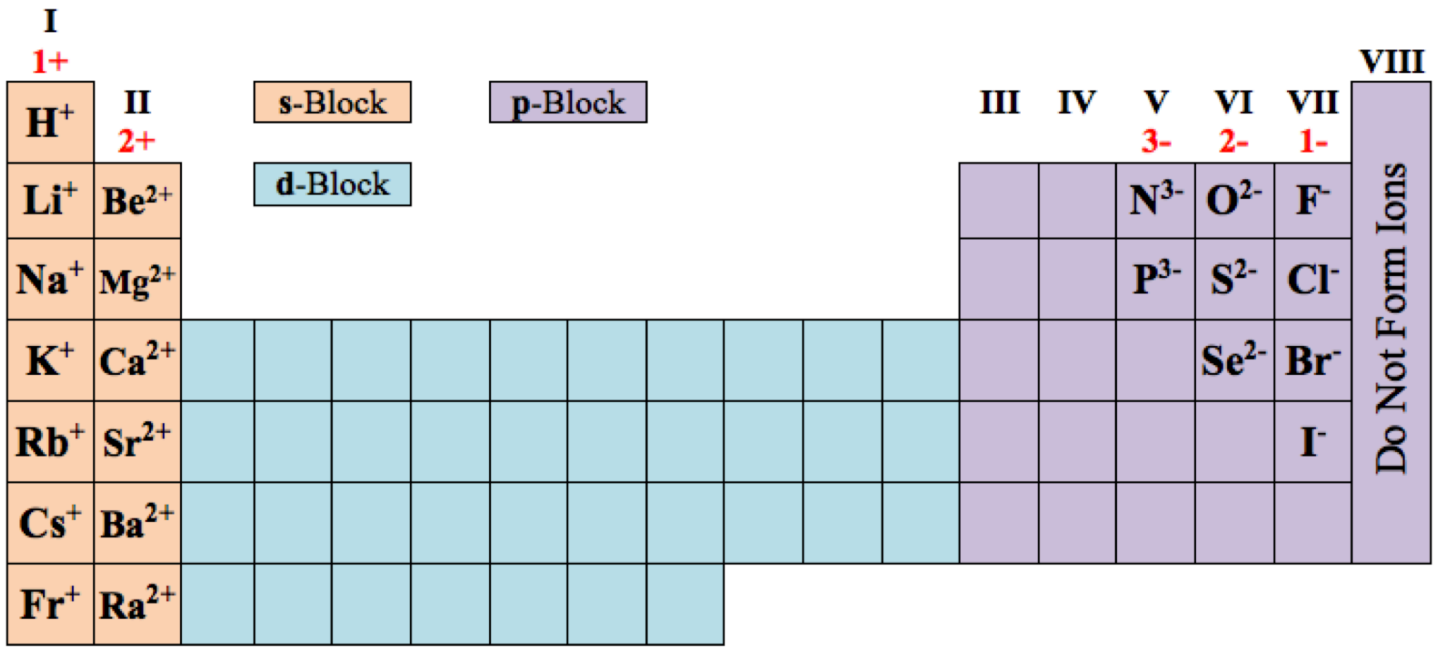
Group VI	6	Gain 2 electrons	2-
Group VII	7	Gain 1 electron	1-
Group VIII	8	Do not form ions, noble gas atoms have filled outer shells.	

EXPLANATION:

Ions generally form such that the ***ion*** has an **octet of electrons** in its outermost shell. This tendency will allow us to predict the charge of the ion that is formed for particular elements.



Go back

For more details: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.

Go to next question

3.10) Redraw the table below and then fill in the missing information for the particular species in each of the column.

Element:	Se		
Mass number:	76	65	
Number of neutrons:		36	39
Number of protons:			36
Number of electrons:			36
charge:	2-	1+	

NOTE: You did a similar problem in chapter 2. In that problem, *all of the species were atoms*. All atoms are electrically neutral (charge = 0) because the number of protons is the same as the number of electrons. In this problem, ***one or more of the species are ions***.

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.10) Redraw the table below and then fill in the missing information for the particular species in each of the column.

Element:	Se		
Mass number:	76	65	
Number of neutrons:		36	39
Number of protons:			36
Number of electrons:			36
charge:	2-	1+	

HINT: You did a similar problem in chapter 2. In that problem, *all of the species were atoms*. All atoms are electrically neutral (total charge = 0) because the number of protons is the same as the number of electrons. In this problem, ***one or more of the species are ions***. Cations have more protons than electrons; this gives them a positive charge. Anions have more electrons that protons; this gives them a negative charge.

The ***mass number*** of an atom is defined as *the number of protons* ***plus*** *the number of neutrons*. Whenever you know the two of the three quantities in the equations shown below, you can determine the value for the third quantity.

mass number = number of protons + number of neutrons

For more help: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.10) Redraw the table below and then fill in the missing information for the particular species in each of the column.

Element:	Se	Cu	Kr
Mass number:	76	65	75
Number of neutrons:	42	36	39
Number of protons:	34	29	36
Number of electrons:	36	28	36
charge:	2-	1+	0

EXPLANATION: All atoms are electrically neutral (total charge = 0) because the number of protons is the same as the number of electrons. In this problem, Kr is atom because it has the same number of protons and electrons.

Cations have more protons than electrons; this gives them a net positive charge. Anions have more electrons that protons; this gives them a net negative charge. In this problem, **Se** ***and*** **Cu** ***are ions***. You were told that **Se** has a **2-** charge, it will therefore have two more electrons than it does protons. You were told that **Cu** has a **1+** charge, it will therefore have one more proton than it does electrons.

The ***mass number*** of an atom is defined as *the number of protons* ***plus*** *the number of neutrons*. Whenever you know the two of the three quantities in the equations shown below, you can determine the value for the third quantity.

mass number = number of protons + number of neutrons

Go back

For more details: See chapter 3 part 5 video or chapter 3 section 4 in the textbook.

Go to next question

3.11) Write the ***formula*** for each of the following compounds:

a) potassium bicarbonate

b) sodium bromide

c) iron(III) fluoride

d) sodium carbonate

e) iron(II) sulfate

f) barium hydroxide

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.11) Write the ***formula*** for each of the following compounds:

a) potassium bicarbonate

b) sodium bromide

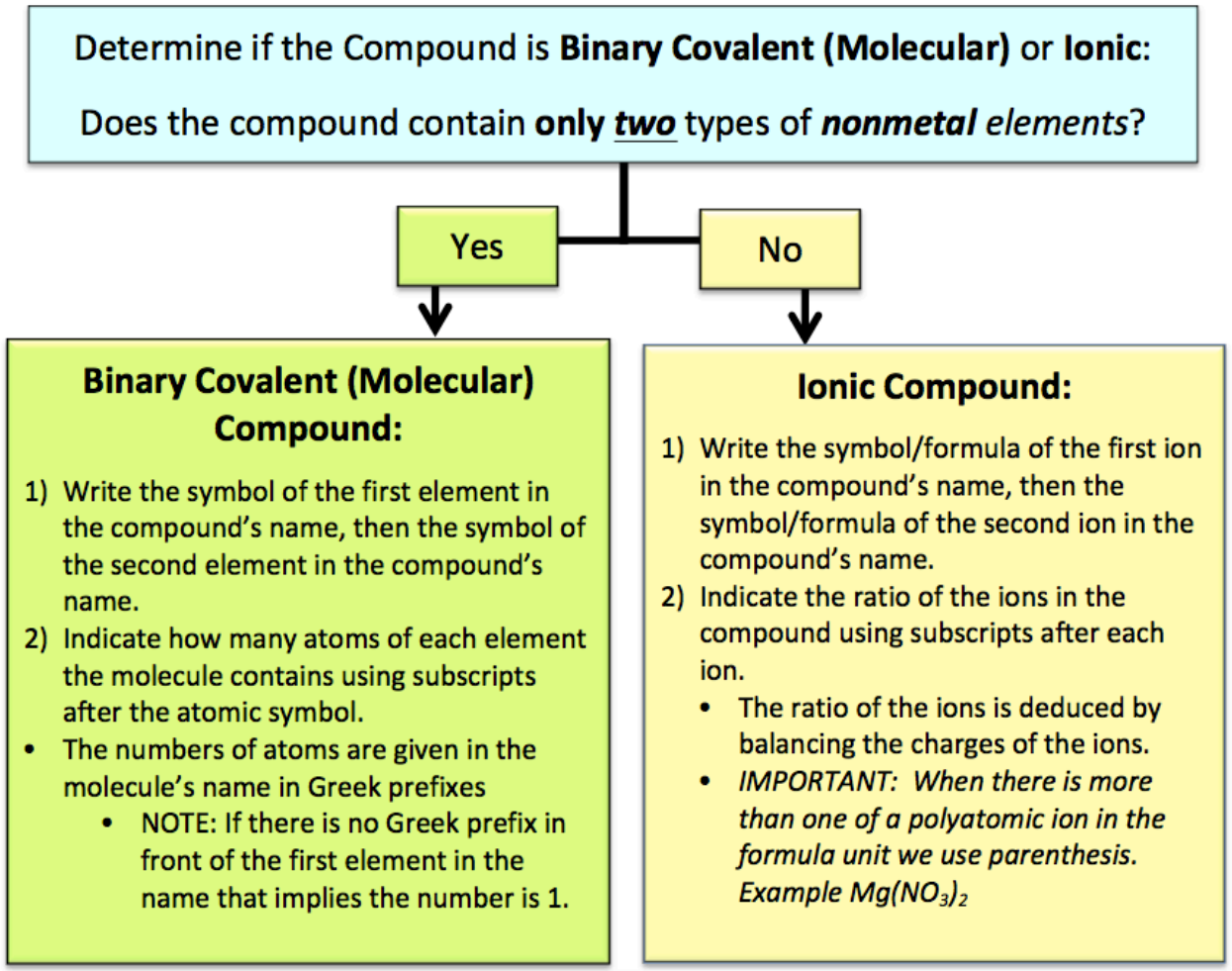
c) iron(III) fluoride

d) sodium carbonate

e) iron(II) sulfate

f) barium hydroxide

HINT: Use the diagram below to find the formula of a compound when you are given its name.



Go back

For more help: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.

Click here to **check your answer**

Go to next question

3.11) Write the ***formula*** for each of the following compounds:

a) potassium bicarbonate **$KHCO_3$**

(Note that bicarbonate is a *polyatomic ion*; see the table of polyatomic ions to get its formula and charge.)

b) sodium bromide **NaBr**

c) iron(III) fluoride **FeF_3**

- From the (III) in its name, we know the iron has a 3+ charge

d) sodium carbonate **Na_2CO_3**

e) iron(II) sulfate **$FeSO_4$**

- We **do not** write **$Fe_2(SO_4)_2$**, we write the lowest ratio, 2:2 is equal to 1:1

f) barium hydroxide **$Ba(OH)_2$**

- Because hydroxide **(OH)** is a **polyatomic ion *and*** its ***subscript*** is ***greater than 1***, we write its formula in **parenthesis** and the subscript is written after/outside of the parenthesis.

Go back

For more details: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.

Go to next question

3.12) Write the names for the following compounds. Be sure to use roman numerals with cations that can occur with multiple charges.

a) $Fe_2(CO_3)_3$

b) $Cu(OH)_2$

c) $(NH_4)_2SO_4$

d) $LiNO_3$

e) $Mg(NO_3)_2$

f) $AgCl$

g) $Al(OH)_3$

h) $CaSO_4$

i) FeS

j) $PbCl_2$

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.12) Write the names for the following compounds. Be sure to use roman numerals with cations that can occur with multiple charges.

a) $Fe_2(CO_3)_3$

b) $Cu(OH)_2$

c) $(NH_4)_2SO_4$
- Note: NH_4^+ is a polyatomic ***cation***

d) $LiNO_3$

e) $Mg(NO_3)_2$

f) $AgCl$

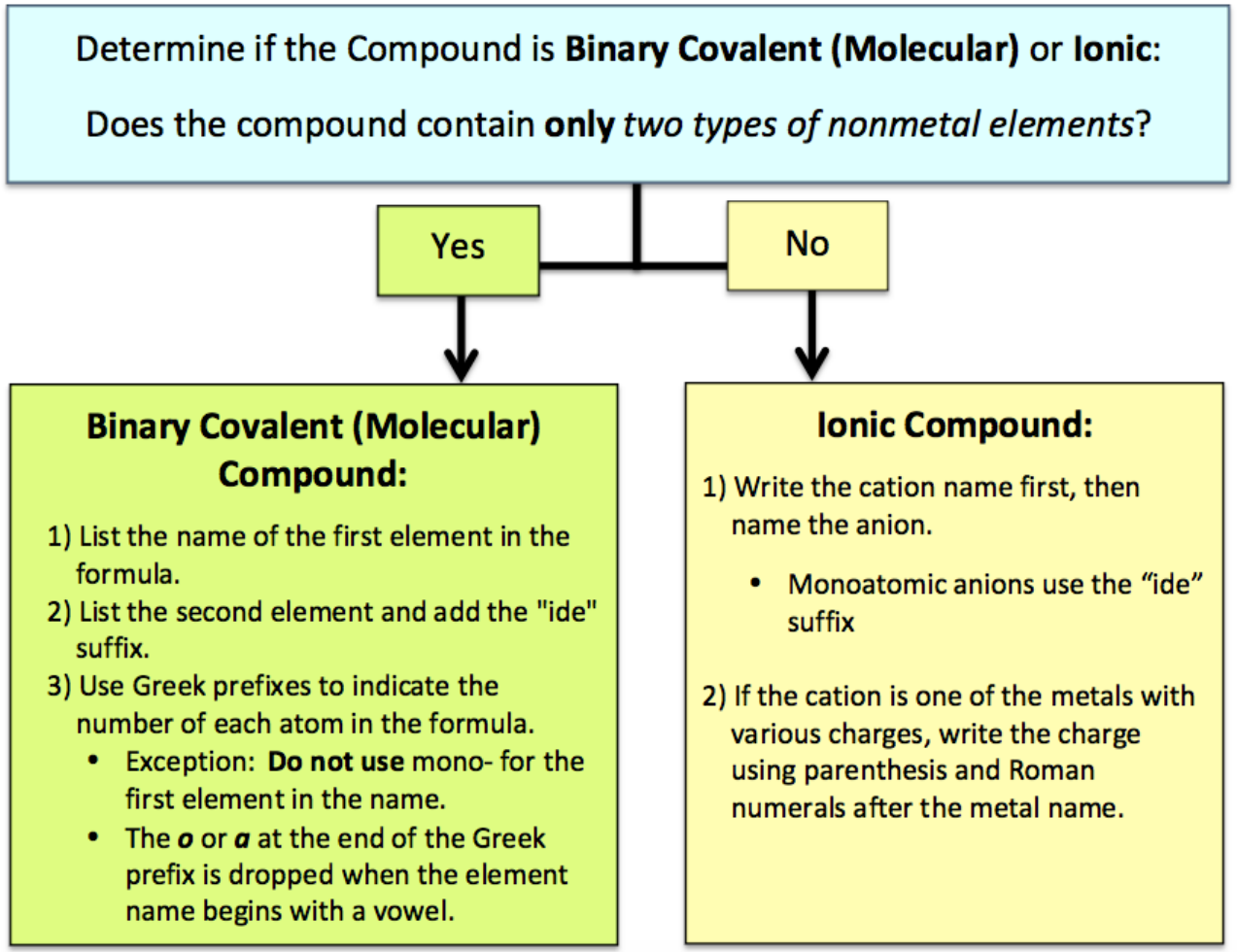
g) $Al(OH)_3$

h) $CaSO_4$

i) FeS

j) $PbCl_2$

HINT: Use the diagram below to find the formula of a compound when you are given its name.



Go back

For more help: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.

Click here to **check your answer**

Go to next question

3.12) Write the names for the following compounds. Be sure to use roman numerals with cations that can occur with multiple charges.

a) $Fe_2(CO_3)_3$ **iron(III) carbonate** Iron (Fe) is one of the metals that can occur with varying charge, you will need to **deduce** the charge on the iron cation based on the three carbonates and their charge.

b) $Cu(OH)_2$ **copper(II) hydroxide** Cu is one of the metals that can occur with varying charge.

c) $(NH_4)_2SO_4$ **ammonium sulfate**

- Note: NH_4^+ is a polyatomic ***cation***

d) $LiNO_3$ **lithium nitrate**

e) $Mg(NO_3)_2$ **magnesium nitrate**

f) $AgCl$ **silver chloride** Silver (Ag) is a transition metal, however it is not one that can occur with varying charge, therefore you do not need to indicate the charge in parenthesis after the cation name.

g) $Al(OH)_3$ **aluminum hydroxide**

h) $CaSO_4$ **calcium sulfate**

i) FeS **iron(II) sulfide** Iron (Fe) is one of the metals that can occur with varying charge.

j) $PbCl_2$ **lead(II) chloride** Lead (Pb) is one of the metals that can occur with varying charge.

Go back

For more details: See chapter 3 part 10 video or chapter 3 section 7 in the textbook.

Go to next question

3.13) Which of the following atom pairs would most likely be connected by a **covalent bond**?

a) Na and O

b) Na and Cl

c) O and Cl

d) K and F

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.13) Which of the following atom pairs would most likely be connected by a **covalent bond**?

a) Na and O

b) Na and Cl

c) O and Cl

d) K and F

HINT:

Covalent bonding results from:

- Combining **nonmetal ions** with **nonmetal ions**.

Ionic bonding (ionic compounds) results from:

- Combining **metal ions** with **nonmetal ions**.

or

- Combining **polyatomic ions** with ***other*** **ions**.

Go back

Click here to **check your answer**

Go to next question

3.13) Which of the following atom pairs would most likely be connected by a **covalent bond**?

a) Na and O

b) Na and Cl

c) O and Cl

d) K and F

EXPLANATION:

Covalent bonding results from:

Combining **nonmetal ions** with **nonmetal ions**.

Ionic bonding (ionic compounds) results from:

- Combining **metal ions** with **nonmetal ions**.

or

- Combining **polyatomic ions** with ***other*** **ions**.

Go back

Go to next question

3.14) Which ion has the same number of *total electrons* as the noble gas argon (Ar)?

a) F^-

b) Br^-

c) Mg^{2+}

d) Ca^{2+}

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.14) Which ion has the same number of *total electrons* as the noble gas argon (Ar)?

a) F^-

b) Br^-

c) Mg^{2+}

d) Ca^{2+}

HINT:

In this problem, we are concerned with the number of **total electrons**, **NOT** the number of valence electrons.

Determine the number of electrons in Argon.

Determine the number of electrons in the ions by starting with the number of electrons that would be contained in the neutral atoms and then adding or subtracting electrons based on the charge.

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.14) Which ion has the same number of *total electrons* as the noble gas **argon** (Ar)?

a) F^- A fluorine **atom** contains 9 electrons. A fluoride **ion** has a charge of **1-** because it contains one extra electron. This give a total of 10 electrons.

b) Br^- A bromine **atom** contains 35 electrons. A bromide **ion** has a charge of **1-** because it contains one extra electron. This give a total of 36 electrons.

c) Mg^{2+} A magnesium **atom** contains 12 electrons. A magnesium **ion** has a charge of **2+** because it contains two fewer electrons (than the atom). This give a total of 10 electrons.

d) Ca^{2+} A calcium **atom** contains 20 electrons. A calcium **ion** has a charge of **2+** because it contains two fewer electrons (than the atom). This give a total of 18 electrons.

EXPLANATION: In this problem, we are concerned with the number of **total electrons**, **NOT** the number of valence electrons.

Argon has a total of **18 electrons**.

The number of electrons in the ions is determined by considering the number of electrons that are contained in the neutral atoms and then adding or subtracting electrons based on the charge.

For more help: See chapter 3 part 3 video or chapter 3 section 2 in the textbook.

Go back

Go to next question

3.15) Write the names of the following compounds:

a) S_2I_4

b) P_5F_8

c) N_2O_4

d) NBr_3

e) N_2O_5

f) PCl_3

g) H_2S

h) N_2O

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.15) Write the names of the following compounds:

a) S_2I_4

b) P_5F_8

c) N_2O_4

d) NBr_3

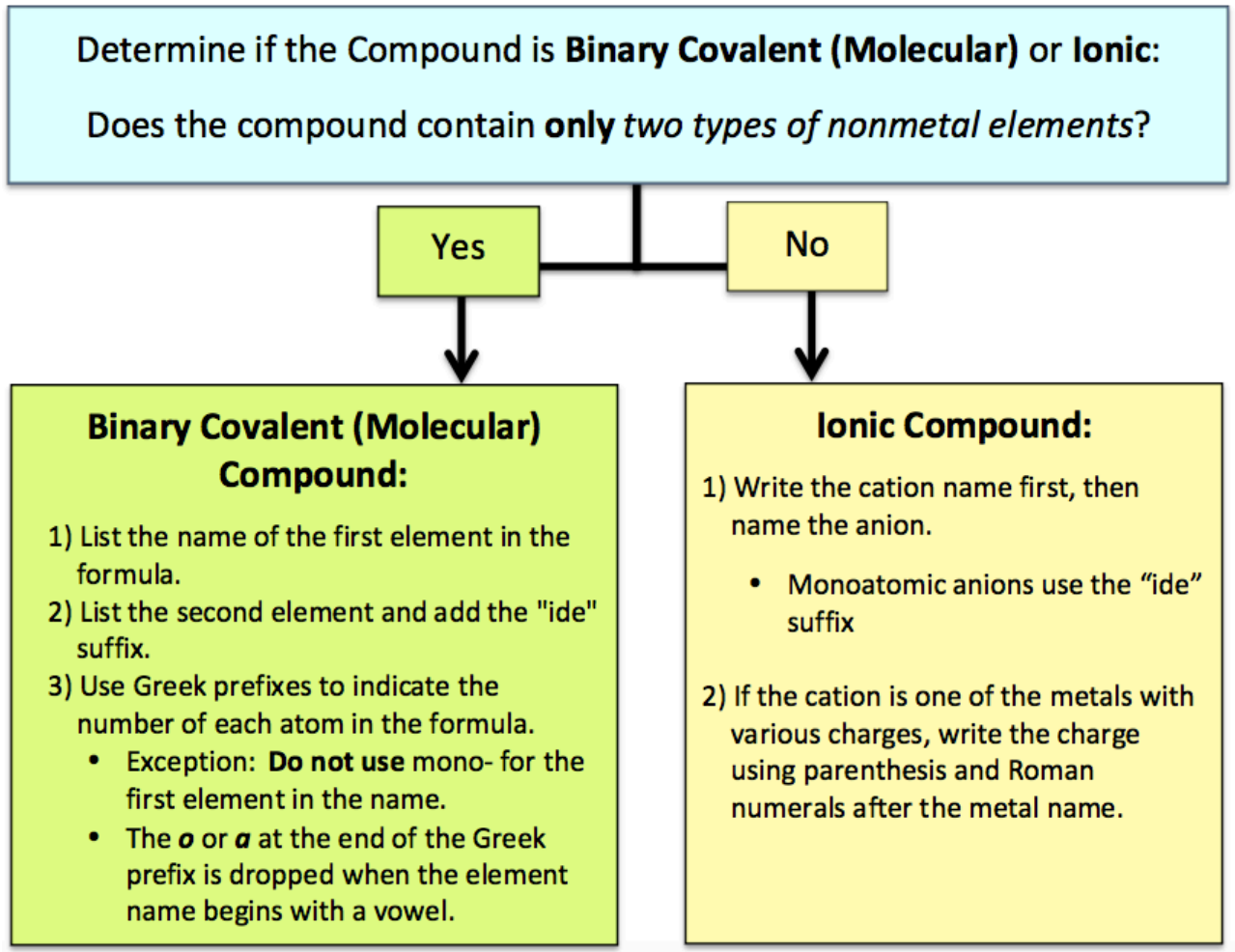
e) N_2O_5

f) PCl_3

g) H_2S

h) N_2O

HINT: Use the diagram below to find the name of a compound when you are given its formula.



For more help: See chapter 3 part 8 video or chapter 3 section 6 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.15) Write the names of the following compounds:

a) S_2I_4 **disulfur tetriodide**

b) P_5F_8 **pentaphosphorus octafluoride**

c) N_2O_4 **dinitrogen tetroxide**

d) NBr_3 **nitrogen tribromide**

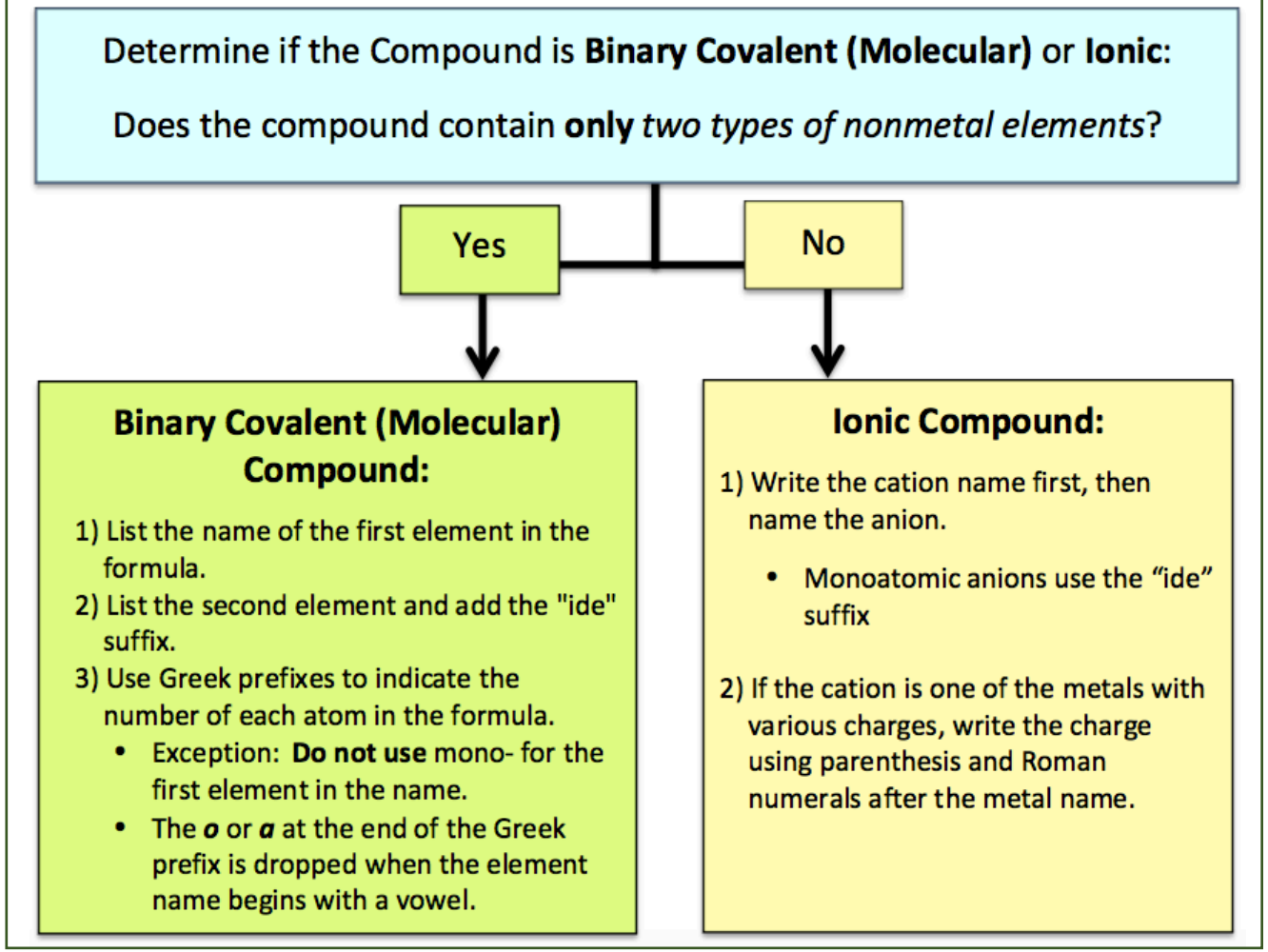
e) N_2O_5 **dinitrogen pentoxide**

f) PCl_3 **phosphorus trichloride**

g) H_2S **dihydrogen monosulfide**

h) N_2O **dinitrogen monoxide**

EXPLANATION: Use the diagram below to find the name of a compound when you are given its formula.



For more details: See chapter 3 part 8 video or chapter 3 section 6 in the textbook.

Go back

Go to next question

3.16) Write the molecular formulas of the following compounds:

a) disulfur tetrafluoride

b) carbon trioxide

c) nitrogen pentoxide

d) nitrogen tribromide

e) dinitrogen heptachloride

f) carbon tetrachloride

g) hydrogen monochloride

h) trihydrogen monophosphide

i) dihydrogen monoxide

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.16) Write the molecular formulas of the following compounds:

a) disulfur tetrafluoride

b) carbon trioxide

c) nitrogen pentoxide

d) nitrogen tribromide

e) dinitrogen heptachloride

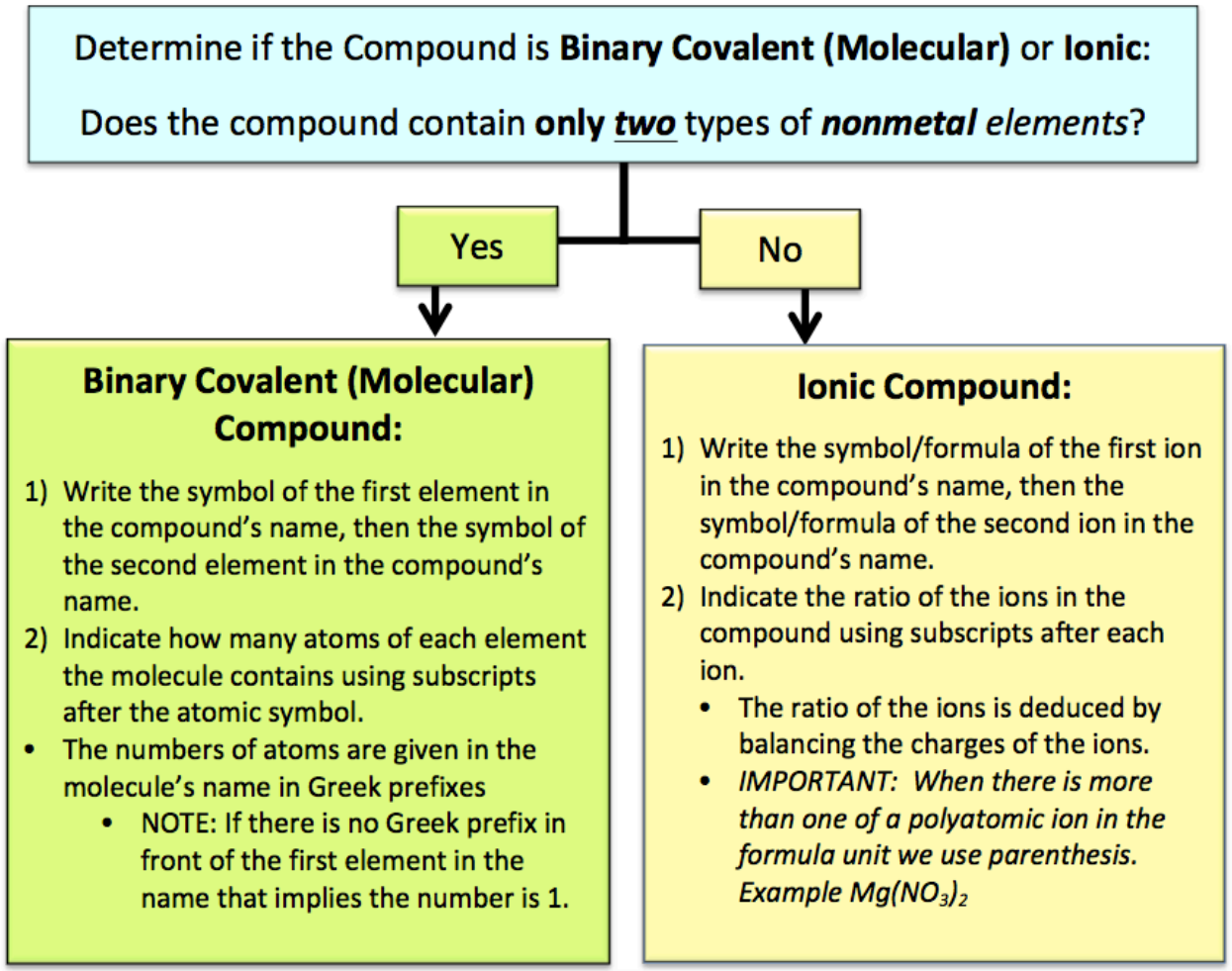
f) carbon tetrachloride

g) hydrogen monochloride

h) trihydrogen monophosphide

i) dihydrogen monoxide

HINT: Use the diagram below to find the formula of a compound when you are given its name.



Go back

For more help: See chapter 3 part 8 video or chapter 3 section 6 in the textbook.

Click here to **check your answer**

Go to next question

3.16) Write the molecular formulas of the following compounds:

a) disulfur tetrafluoride S_2F_4

b) carbon trioxide CO_3

c) nitrogen pentoxide NO_5

d) nitrogen tribromide NBr_3

e) dinitrogen heptachloride N_2Cl_7

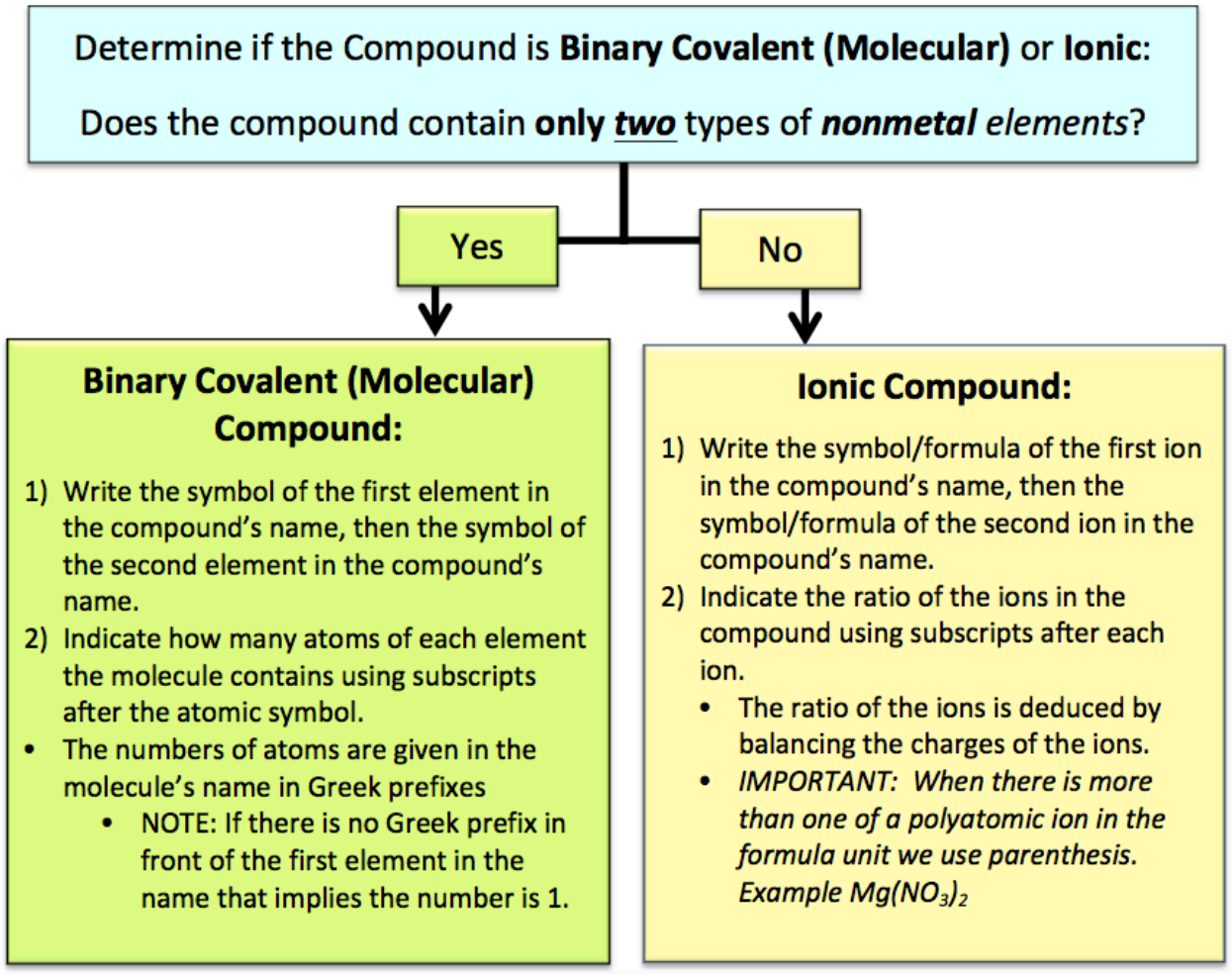
f) carbon tetrachloride CCl_4

g) hydrogen monochloride HCl

h) trihydrogen monophosphide H_3P

i) dihydrogen monoxide H_2O

EXPLANATION: Use the diagram below to find the formula of a compound when you are given its name.



Go back

For more details: See chapter 3 part 8 video or chapter 3 section 6 in the textbook.

Go to next question

3.17) Calculate the molar mass of each compound given below.

REMINDER: In order for your answers to ***exactly*** match the solutions provided for review problems, round ***molar masses*** to **two digits** *to the right of the decimal point* when possible.

a) H_2S

b) iron(II) nitrate

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

3.17) Calculate the molar mass of each compound given below.

REMINDER: In order for your answers to ***exactly*** match the solutions provided for review problems, round ***molar masses*** to **two digits** *to the right of the decimal point* when possible.

a) H_2S

HINT: We cannot get the molar mass of a compound *directly* from the periodic table as we did for atomic molar masses.

To calculate the molar mass of a **compound** we **add up** the ***atomic molar masses*** of **all** atoms in the chemical formula.

b) iron(II) nitrate

For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.17) Calculate the molar mass of each compound given below.

a) H_2S **34.09 g/mole**

CLICK HERE to see the **complete solution** for this problem

b) iron(II) nitrate **179.87 g/mole**

Go back

Go to next question

3.17) Calculate the molar mass of each compound given below.

REMINDERS: In order for your answers to ***exactly*** match the solutions provided for review problems, round ***molar masses*** to **two digits** *to the right of the decimal point* when possible.

a) H_2S **34.09 g/mole**

Atom	# of Atoms	Atomic Molar Mass	Total
sulfur	1	x 32.07 g/mole	32.07 g/mole
hydrogen	2	x 1.01 g/mole	2.02 g/mole
		Molar Mass of H_2S =	**34.09 g/mole**

b) iron(II) nitrate **$Fe(NO_3)_2$** **179.87 g/mole**

Fe — ***one*** **mole** of *iron*(II) *ions*

$(NO_3)_2$ — ***two*** **moles** of *nitrate ions* contain:

- ***two*** moles of nitrogen
- ***six*** (2 x 3) moles of oxygen

one mole of **Fe**: **1** x 55.85 g/mole = **55.85** g/mole
two moles of **N**: **2** x 14.01 g/mole = **28.02** g/mole
six moles of **O**: **6** x 16.00 g/mole = **96.00** g/mole

The molar mass of $Fe(NO_3)_2$ is **179.87** g/mole

Go back

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go to next question

3.18) How many moles are contained in 251 grams of Fe_2O_3?

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

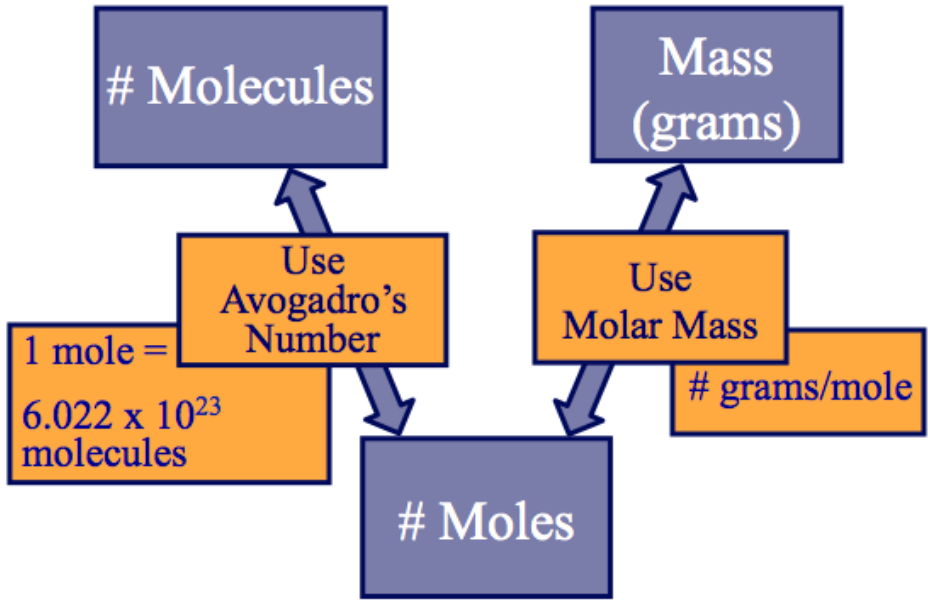
3.18) How many moles are contained in 251 grams of Fe_2O_3?

HINT:

First determine the **molar mass** of Fe_2O_3.

Next, convert from grams to moles using the molar mass as a **conversion factor**.

The chart shown below can be helpful when doing gram/mole/molecule conversions.



For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.18) How many moles are contained in 251 grams of Fe_2O_3?

ANSWER: 1.57 moles

Your answer should have ***three*** significant figures.

CLICK HERE to see the **complete solution** for this problem

Go back

Go to next question

3.18) How many moles are contained in 251 grams of Fe_2O_3?

ANSWER: 1.57 moles

Your answer should have ***three*** significant figures.

First determine the **molar mass** of Fe_2O_3.

Ion	# of ions in the Formula Unit	Molar Mass of ion	Total
iron (III)	2	x 55.85 g/mole	= 111.70 g/mole
oxide	3	x 16.00 g/mole	= 48.00 g/mole
Molar Mass (Formula Mass) of Fe_2O_3			**= 159.70 g/mole**

Next, convert from grams to moles using the molar mass as a ***conversion factor***.

The relationship between ***moles*** and and ***grams*** is the ***molar mass***.

$$251 \text{ grams } Fe_2O_3 \times \frac{1 \text{ mole } Fe_2O_3}{159.70 \text{ grams } Fe_2O_3} = \mathbf{1.57 \text{ moles } Fe_2O_3}$$

Go back

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go to next question

3.19) How many grams are contained in 0.482 moles of Na_2CO_3?

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

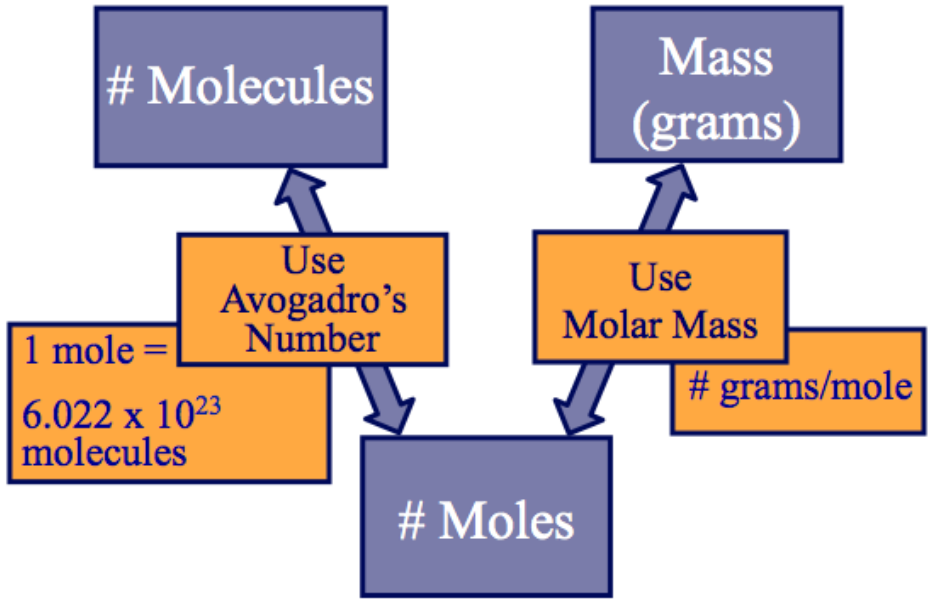
3.19) How many grams are contained in 0.482 moles of Na_2CO_3?

HINT:

First determine the **molar mass** of Na_2CO_3.

Next, convert from moles to grams using the molar mass as a **conversion factor**.

The chart shown below can be helpful when doing gram/mole/molecule conversions.



For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.19) How many grams are contained in 0.482 moles of Na_2CO_3?

ANSWER: 51.1 grams

Your answer should have ***three*** significant figures.

CLICK HERE to see the **complete solution** for this problem

Go back

Go to next question

3.19) How many grams are contained in 0.482 moles of Na_2CO_3?

ANSWER: 51.1 grams

Your answer should have ***three*** significant figures.

First determine the **molar mass** of Na_2CO_3.

Na

two **moles** of *sodium ions*

(CO_3)

one **mole** of *carbonate ions* contains:

- ***one*** mole of carbon
- ***three*** moles of oxygen

two moles of **Na**:	**2** x 22.99 g/mole	=	**45.98** g/mole
one mole of **C**:	**1** x 12.01 g/mole	=	**12.01** g/mole
three moles of **O**:	**3** x 16.00 g/mole	= +	**48.00** g/mole
The molar mass of Na_2CO_3 is			**105.99** g/mole

Next, convert from moles to grams using the molar mass as a ***conversion factor***.

The relationship between ***moles*** and and ***grams*** is the ***molar mass***.

$$0.482 \text{ moles } Na_2CO_3 \times \frac{105.99 \text{ grams } Na_2CO_3}{1 \text{ mole } Na_2CO_3} = \mathbf{51.1 \text{ grams } Na_2CO_3}$$

Go back

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go to next question

3.20) What is the mass (in grams) of 9.39×10^{24} **molecules** of methanol (CH_4O)?

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

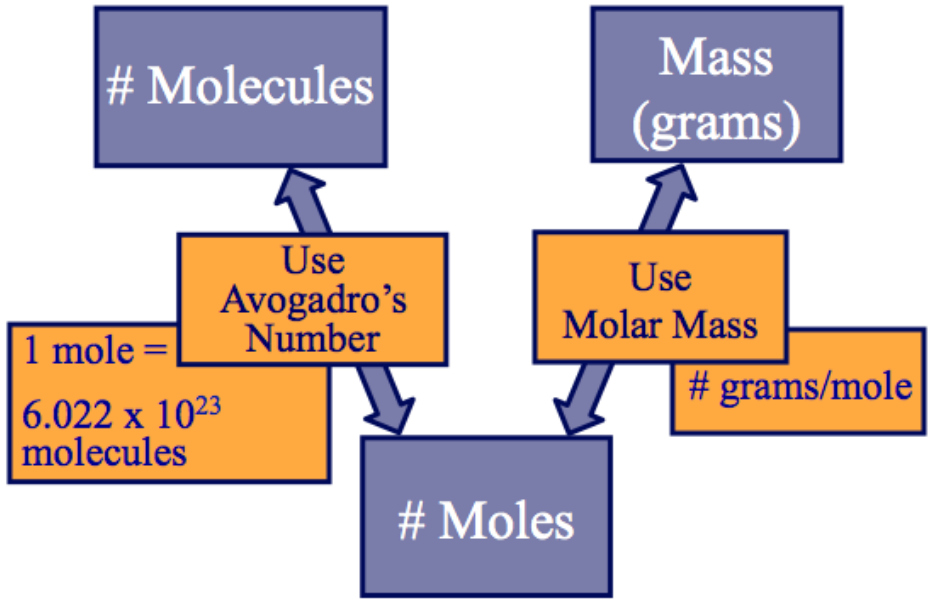
3.20) What is the mass (in grams) of 9.39×10^{24} **molecules** of methanol (CH_4O)?

HINT:

First determine the **molar mass** of CH_4O.

Next, convert from **molecules** to **grams**. This is a **TWO STEP** *unit conversion problem*.

The chart shown below can be helpful when doing gram/mole/molecule conversions.



For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.20) What is the mass (in grams) of 9.39×10^{24} **molecules** of methanol (CH_4O)?

ANSWER: 500. grams CH_4O
(or 5.00×10^2 grams CH_4O)

Your answer should have ***three*** significant figures.

CLICK HERE to see the **complete solution** for this problem

Go back

Go to next question

3.20) What is the mass (in grams) of 9.39×10^{24} **molecules** of methanol (CH_4O)?

ANSWER: 500. grams CH_4O (**or** 5.00×10^2 grams CH_4O)

Your answer should have ***three*** significant figures.

First determine the **molar mass** of CH_4O.

one mole of **C**:	**1** x 12.01 g/mole	=	**12.01** g/mole
four moles of **H**:	**4** x 1.01 g/mole	=	**4.04** g/mole
one mole of **O**:	**1** x 16.00 g/mole	= +	**16.00** g/mole
		The molar mass of CH_4O is	**32.05** g/mole

Next, convert from **molecules** to **grams**. This is a **TWO STEP** *unit conversion problem*.

First step: convert from ***molecules*** to ***moles***:

$$\frac{9.39 \times 10^{24}\ CH_4O\ \text{molecules}}{\left|\begin{array}{c} 1\ \text{mole}\ CH_4O \\ \hline 6.022 \times 10^{23}\ CH_4O\ \text{molecules} \end{array}\right.} = 15.5928263\ CH_4O\ \text{molecules (unrounded)}$$

Second step: convert from ***moles*** to ***grams***:

$$\frac{15.5928263\ CH_4O\ \text{molecules}}{\left|\begin{array}{c} 32.05\ \text{grams}\ CH_4O \\ \hline 1\ \text{mole}\ CH_4O \end{array}\right.} = 500.\ \text{grams}\ CH_4O\ (\textbf{or}\ 5.00 \times 10^2\ \text{grams}\ CH_4O)$$

Alternatively, these two steps can be combined into one equation:

$$\frac{9.39 \times 10^{24}\ CH_4O\ \text{molecules}}{\left|\begin{array}{c} 1\ \text{mole}\ CH_4O \\ \hline 6.022 \times 10^{23}\ CH_4O\ \text{molecules} \end{array}\right.} \times \frac{32.05\ \text{grams}\ CH_4O}{1\ \text{mole}\ CH_4O} = 500.\ \text{grams}\ CH_4O\ (\textbf{or}\ 5.00 \times 10^2\ \text{grams}\ CH_4O)$$

Go back

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go to next question

3.21) What is the mass (in grams) of 9.43×10^{23} molecules of sucrose ($C_{12}H_{22}O_{11}$)?

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

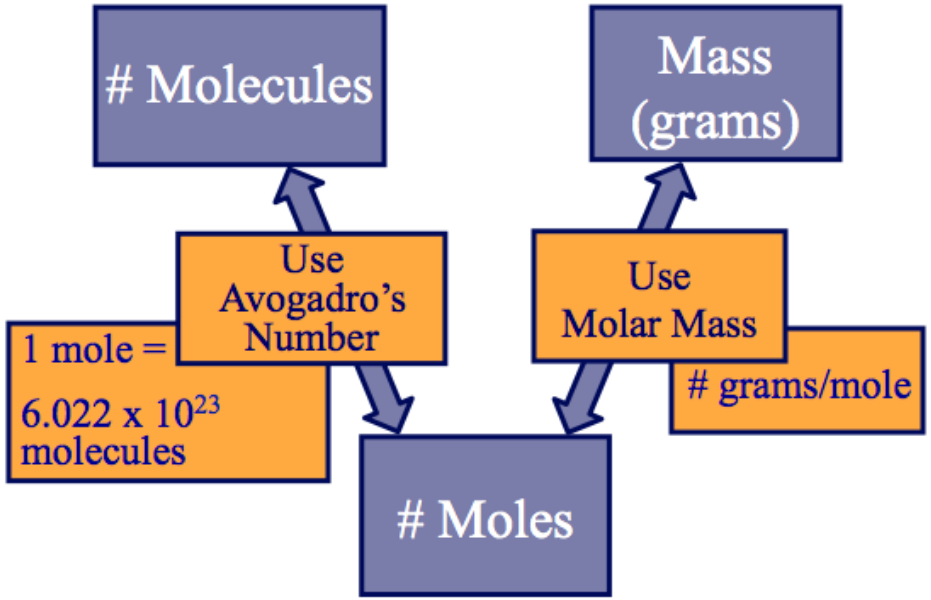
3.21) What is the mass (in grams) of 9.43×10^{23} molecules of sucrose ($C_{12}H_{22}O_{11}$)?

HINT:

First determine the **molar mass** of $C_{12}H_{22}O_{11}$.

Next, convert from **molecules** to **grams**. This is a **TWO STEP** *unit conversion problem*.

The chart shown below can be helpful when doing gram/mole/molecule conversions.



For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.21) What is the mass (in grams) of 9.43×10^{23} molecules of sucrose ($C_{12}H_{22}O_{11}$)?

ANSWER: 536 grams $C_{12}H_{22}O_{11}$

Your answer should have ***three*** significant figures.

CLICK HERE to see the **complete solution** for this problem

Go back

Go to next question

3.21) What is the mass (in grams) of 9.43×10^{23} molecules of sucrose ($C_{12}H_{22}O_{11}$)?

ANSWER: 536 grams $C_{12}H_{22}O_{11}$

Your answer should have ***three*** significant figures.

First determine the **molar mass** of $C_{12}H_{22}O_{11}$.

12 moles of **C**:	**12** x 12.01 g/mole =	**144.12** g/mole
22 moles of **H**:	**22** x 1.01 g/mole =	**22.22** g/mole
11 moles of **O**:	**11** x 16.00 g/mole =	+ **176.00** g/mole
The molar mass of $C_{12}H_{22}O_{11}$ is		**342.34** g/mole

Next, convert from **molecules** to **grams**. This is a **TWO STEP** *unit conversion problem*.

First step: convert from ***molecules*** to ***moles***:

$$\frac{9.43 \times 10^{23}\ C_{12}H_{22}O_{11}\ \text{molecules} \quad | \quad 1\ \text{mole}\ C_{12}H_{22}O_{11}}{6.022 \times 10^{23}\ C_{12}H_{22}O_{11}\ \text{molecules}} = 1.565924942\ C_{12}H_{22}O_{11}\ \text{molecules}$$

(unrounded)

Second step: convert from ***moles*** to ***grams***:

$$\frac{1.565924942\ C_{12}H_{22}O_{11}\ \text{molecules} \quad | \quad 342.34\ \text{grams}\ C_{12}H_{22}O_{11}}{1\ \text{mole}\ C_{12}H_{22}O_{11}} = 536\ \text{grams}\ C_{12}H_{22}O_{11}$$

Alternatively, these two steps can be combined into one equation:

$$\frac{9.43 \times 10^{23}\ C_{12}H_{22}O_{11}\ \text{molecules} \quad | \quad 1\ \text{mole}\ C_{12}H_{22}O_{11} \quad | \quad 342.34\ \text{grams}\ C_{12}H_{22}O_{11}}{6.022 \times 10^{23}\ C_{12}H_{22}O_{11}\ \text{molecules} \quad | \quad 1\ \text{mole}\ C_{12}H_{22}O_{11}} = 536\ \text{grams}\ C_{12}H_{22}O_{11}$$

Go back

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go to next question

3.22) How many molecules are contained in 38.33 grams of NO_2?

Go back

Click here for a **hint**

Click here to **check your answer**

Go to next question

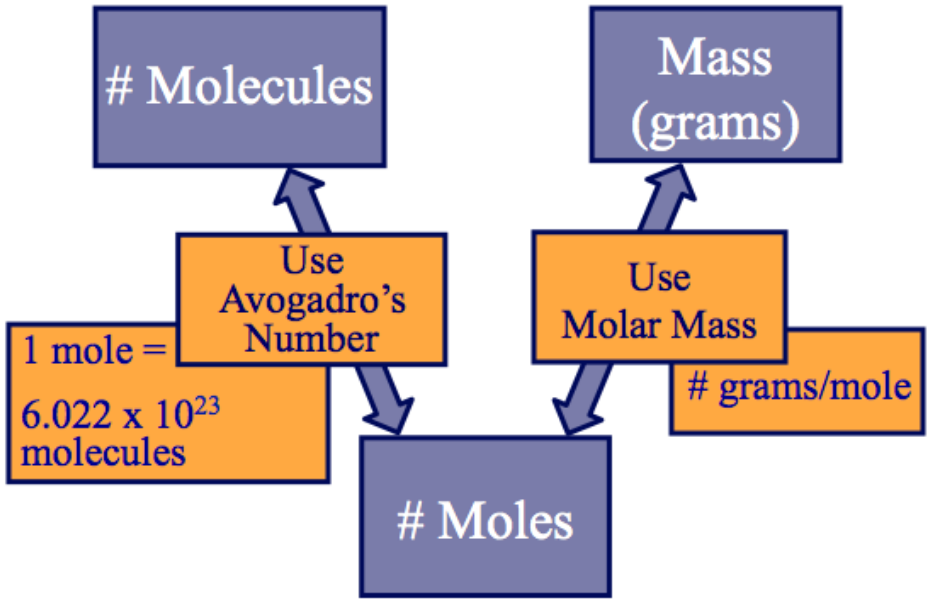
3.22) How many molecules are contained in 38.33 grams of NO_2?

HINT:

First determine the **molar mass** of NO_2.

Next, convert from **grams** to **molecules**. This is a **TWO STEP** *unit conversion problem*.

The chart shown below can be helpful when doing gram/mole/molecule conversions.



For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

Go to next question

3.22) How many molecules are contained in 38.33 grams of NO_2?

ANSWER: 5.017 x 10^{23} NO_2 molecules

Your answer should have ***four*** significant figures.

CLICK HERE to see the **complete solution** for this problem

Go back

Go to next question

3.22) How many molecules are contained in 38.33 grams of NO_2?

ANSWER: 5.017 x 10^{23} NO_2 molecules

Your answer should have ***four*** significant figures.

First determine the **molar mass** of NO_2.

one mole of **N**:	1 x 14.01 g/mole =	**14.01** g/mole
two moles of **O**:	2 x 16.00 g/mole = ✚	**32.00** g/mole
	The molar mass of NO_2 is	**46.01** g/mole

Next, convert from **grams** to **molecules**. This is a **TWO STEP** *unit conversion problem*.

First step: convert from ***grams*** to ***moles***:

$$\frac{38.33 \text{ grams } NO_2}{46.01 \text{ grams } NO_2} \left| \frac{1 \text{ mole } NO_2}{} \right. = 0.8330797653 \text{ moles } NO_2 \text{ (unrounded)}$$

Second step: convert from ***moles*** to ***molecules***:

$$\frac{0.8330797653 \text{ moles } NO_2}{1 \text{ mole } NO_2} \left| \frac{6.022 \times 10^{23} \text{ } NO_2 \text{ molecules}}{} \right. = \mathbf{5.017 \times 10^{23} \text{ } NO_2 \text{ molecules}}$$

Alternatively, these two steps can be combined into one equation:

$$\frac{38.33 \text{ grams } NO_2}{46.01 \text{ grams } NO_2} \left| \frac{1 \text{ mole } NO_2}{1 \text{ mole } NO_2} \right| \frac{6.022 \times 10^{23} \text{ } NO_2 \text{ molecules}}{} = \mathbf{5.017 \times 10^{23} \text{ } NO_2 \text{ molecules}}$$

Go back

For more details: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go to next question

3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H_2S)?

Go back

Click here for a **hint**

Click here to **check your answer**

This is the last chapter 3 review problem

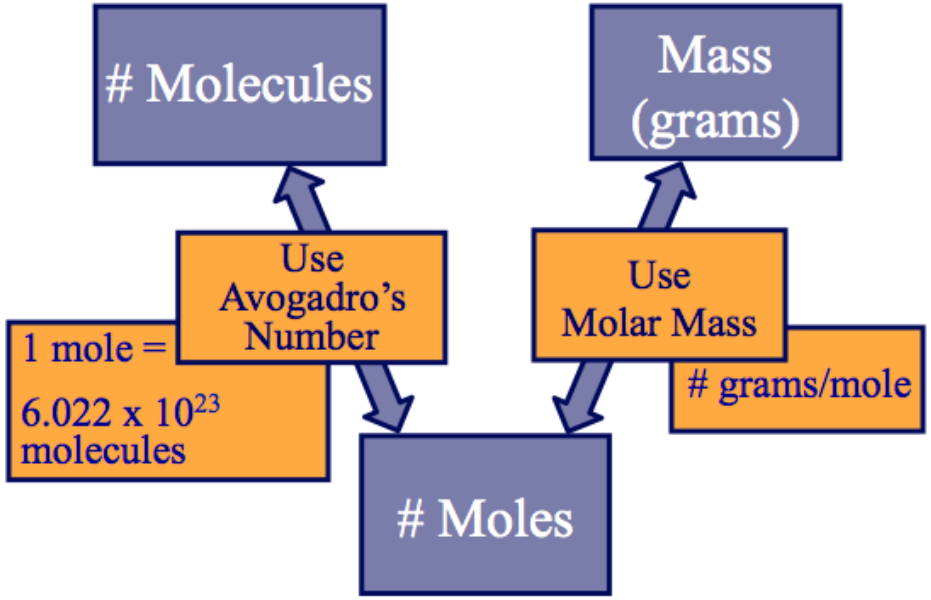
3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H_2S)?

HINT:

First determine the **molar mass** of NO_2.

Next, convert from **grams** to **molecules**. This is a **TWO STEP** *unit conversion problem*.

The chart shown below can be helpful when doing gram/mole/molecule conversions.



For more help: See chapter 3 part 11 video or chapter 3 section 9 in the textbook.

Go back

Click here to **check your answer**

This is the last chapter 3 review problem

3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H_2S)? **ANSWER: 4.457 x 10^{23} H_2S molecules**

Your answer should have ***four*** significant figures.

CLICK HERE to see the **complete solution** for this problem

Go back

This is the last chapter 3 review problem

3.23) How many molecules are contained in 25.23 grams of dihydrogen sulfide (H_2S)? **ANSWER: 4.457 x 10^{23} H_2S molecules**

Your answer should have ***four*** significant figures.

First determine the **molar mass** of H_2S.

two moles of **H**:	2 x 1.01 g/mole =	**2.02** g/mole
one mole of **S**:	1 x 32.07 g/mole = +	**32.07** g/mole
The molar mass of H_2S is		**34.09** g/mole

Next, convert from **grams** to **molecules**. This is a **TWO STEP** *unit conversion problem*.

First step: convert from ***grams*** to ***moles***:

$$\frac{25.23 \text{ grams } H_2S}{} \left| \frac{1 \text{ mole } H_2S}{34.09 \text{ grams } H_2S} \right. = 0.740099736 \text{ moles } H_2S \text{ (unrounded)}$$

Second step: convert from ***moles*** to ***molecules***:

$$\frac{0.740099736 \text{ moles } H_2S}{} \left| \frac{6.022 \times 10^{23} \text{ } H_2S \text{ molecules}}{1 \text{ mole } H_2S} \right. = \mathbf{4.457 \times 10^{23} \ H_2S \text{ molecules}}$$

Alternatively, these two steps can be combined into one equation:

$$\frac{25.23 \text{ grams } H_2S}{} \left| \frac{1 \text{ mole } H_2S}{34.09 \text{ grams } H_2S} \right| \frac{6.022 \times 10^{23} \text{ } H_2S \text{ molecules}}{1 \text{ mole } H_2S} = \mathbf{4.457 \times 10^{23} \ H_2S \text{ molecules}}$$

Go back

This is the last chapter 3 review problem.