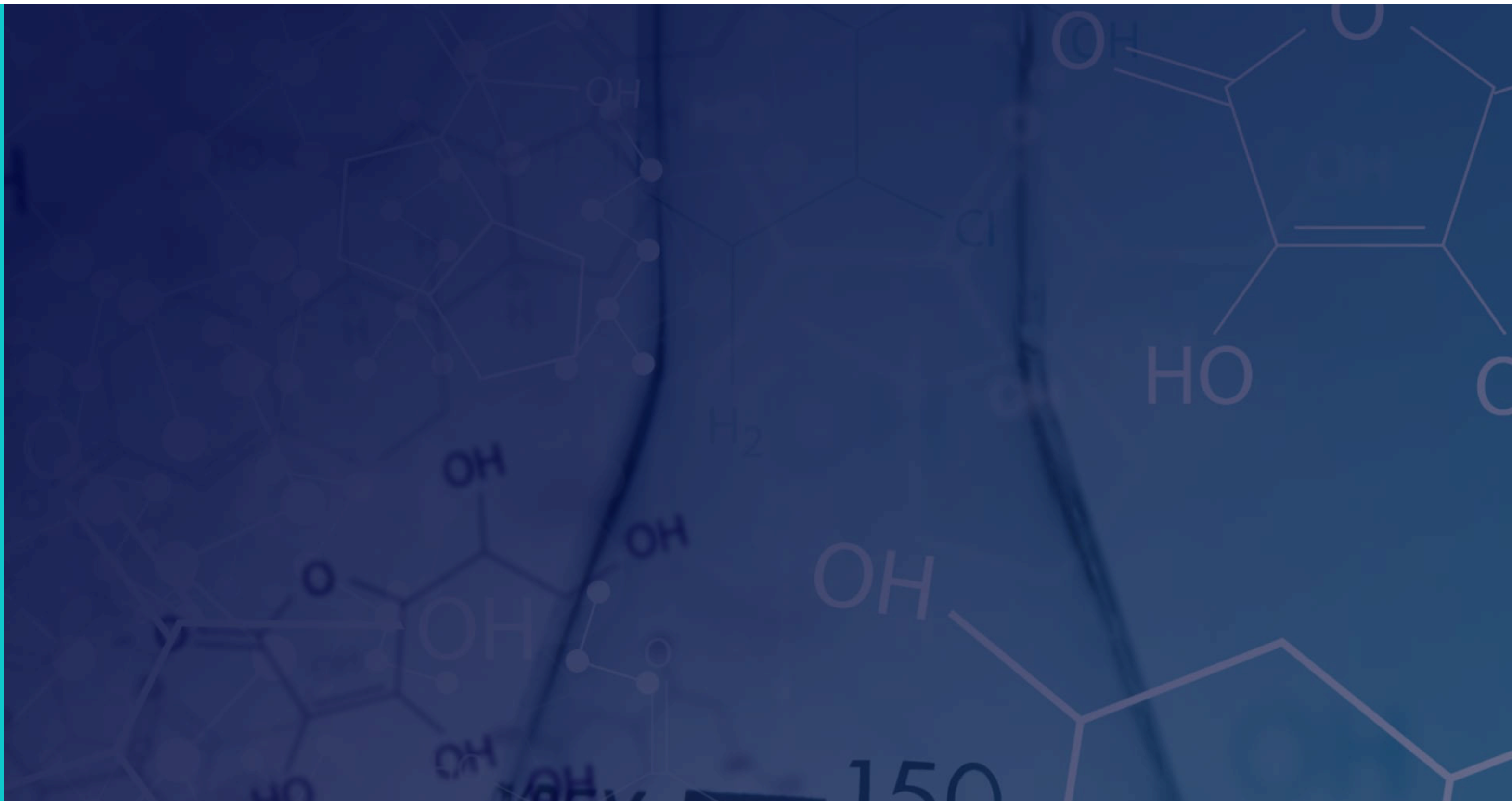


AP Exam Prep

AP Chemistry Exam Prep Course

Summer 2024, Chapter 1 Notes

Table of Contents

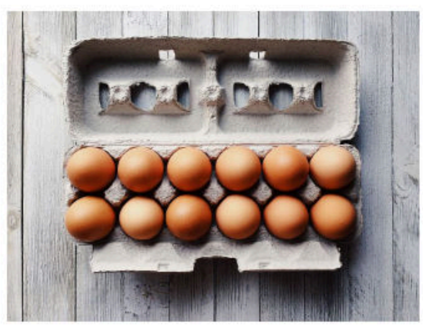
1. Atomic Structure & Properties

1.1 Moles and Molar Mass

1.1.1

Moles

When someone says they want a dozen donuts, roses, or eggs we know they mean they want 12. A **mole** in chemistry also tells us the amount of something.



- Just like how a dozen of something = 12, 1 mole of something =6.02×10^{23} molecules
- 6.02×10^{23} is referred to as **Avagadro's number (N_A)**

WIZE TIP

You should memorize Avagadro's number since it is not always provided on exams! We'll soon see that you need to know it to solve problems :)

$$\boxed{\mathbf{N_A = 6.02 \times 10^{23}}\ \text{molecules/mole}}$$

We use the unit "**moles**" to help make very small amounts more measurable. A mole is just like any other unit!

There are 2 Equations Related to Moles:

$$\boxed{n = \frac{m}{M}} \qquad \boxed{n = \frac{N}{N_A}}$$

n=*# of* ***moles***

m=*mass (****g****)*

M=*molar mass (****g/mol****)*

N=*# of* ***molecules or atoms***

N_A=*Avagadro's number=*6.02×10^{23} ***molecules***

Example #1:

If we are told that a sample of CO_2(s) weighs 11g, how many moles are present?

Example #2

We have 2 moles of CO_2 present in our sample, how many molecules are there?

Example #3

How many oxygen atoms there in 1 mole of CO_2?

1) Find Number of Molecules In the Sample

2) Find the Number of O atoms in the Sample

In each molecule there are ______ O atoms.

Therefore to find the number of O atoms:

# of molecules in sample x ______ O atoms/molecule =# of O atoms

# of O atoms= ________________________________

1.1.2

Example: Calculate the Number of Molecules of Ethyl Mercaptan

The volatile liquid ethyl mercaptan, C_2H_6S, is one of the most odoriferous substances known. It is added to natural gas to make gas leaks detectable. (***d*** = 0.84 g/mL; **MW** = 62.1 g/mol)

1. How many C_2H_6S, molecules are contained in a 3.0 μL sample?

2. In the same 3.0 μL sample, how many C and H atoms are there?

Practice: Converting Mass to Number of Atoms

Calculate the number of nitrogen atoms in 2.25 g of bismuth(III) nitrate.

a) 1.03×10^{22} atoms

b) 1.03×10^{21} atoms

c) 3.43×10^{21} atoms

d) 3.43×10^{22} atoms

Practice: Finding the Number of Moles of Iron

Calculate the number of moles of iron atoms in 14.1 g of iron oxide, Fe_2O_3

A) 0.177 mol

B) 0.0821 mol

C) 0.0906 mol

D) 0.0451 mol

1.2 Mass Spectroscopy of Elements (& Isotopes)

1.2.1 Isotopes and Atomic Mass

Isotopes and Atomic Mass

Isotopes

- When two atoms have the **same number of protons and electrons**, but a **different number of neutrons**, we call these **isotopes.**
- Isotopes have the **same atomic number**, but a **different mass number.**

^{12}C	^{13}C	^{14}C
protons	protons	protons
neutrons	neutrons	neutrons
98 %	1.07 %	Trace

How is the Atomic Mass in the Periodic Table Calculated?

6	12.011
C	
Carbon	

- The atomic mass is written as 12amu for C for example, but this does not mean that each C atom weighs exactly 12amu!
- **The atomic mass # is actually a weighted average** based on the **relative abundance of isotopes**
  - Isotopes have similar reactivity to one another, that's why we can form C bonds with either C-12 or 13 etc

To determine the average mass of an element, use this equation:

$$A.W. = \sum_{i=1}^{n} (mass\ of\ each\ isotope_i)(abundance\ of\ each\ isotope_i)$$

WATCH OUT!

Plug in the **mass of each isotope in amu** and plug in the **relative abundance of each isotope in the form of a decimal** not percentage!

For example, above we are told the relative abundance of C-12 is 98%. You would want to plug in 0.98 for this isotope's relative abundance.

1.2.2

Example: Solving for the Atomic Weight

Chlorine can be found in nature as $^{35}\mathrm{Cl}$ (mass 34.969u, 75.78% abundance) and $^{37}\mathrm{Cl}$ (mass 36.966u, 24.22% abundance). What is the average atomic mass of Cl?

1.2.3

Practice: Solving for the Weight of an Isotope

Naturally occurring potassium contains two stable isotopes. The lighter isotope, ^{39}K (38.9637 amu) is the more abundant isotope, accounting for 93.26% of the nuclei. What is the weight of the heavier isotope, ^{41}K?

- 41.00 amu
- 40.96 amu
- 39.09 amu
- 41.08 amu

1.2.4

Bromine is a noxious fuming red liquid in its elemental form. Naturally occurring bromine has two stable isotopes ^{79}Br (78.918*u*) and ^{81}Br (80.916*u*). What are the relative abundances of these two isotopes?

Enter your answer as a percentage (ex. If the answer is 16.26%, enter 16.26 as the answer)

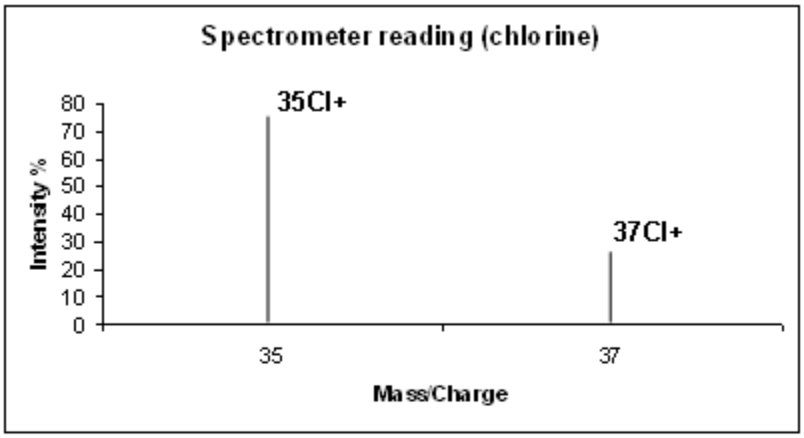
Relative Abundance of Br-79

Relative Abundance of Br-81

1.2.5

Isotopes and the Mass Spectrometer

We get information on isotopes by using a mass spectrometer which gives us a plot like this:



The **height of the peak** tells you the **relative abundance (%).**

Example: ~75% of Cl is Cl-35. ~25% of Cl is Cl-37!

The exact **positions on the x axis** gives you the **mass of each isotope**

Example: There is a peak at x=35, which tells us that there is an isotope of Cl that has a mass of approximately 35.

The **number of peaks** tells us **how many stable isotopes** there are.

Example: Here there are 2 peaks so there are 2 stable isotopes of Cl

In reality, Cl-35 has an abundance of 75.77% and has a mass of 34.96885u.

1.3 Elemental Composition of Pure Substances

1.3.1

Empirical Vs Molecular Formulas

Molecular Formulas

This is the type of formula we are most familiar with. It tells us **exactly how many atoms make up a molecule.**

Example:

C_6H_6 tells us that for each molecule of C_6H_6, there are _____ C atoms and _____ H atoms.

Empirical Formulas

If we know the molecular formula, we can find the empirical formula by dividing the number of atoms by the greatest common factor.

You can think of the empirical formula as the **smallest possible "unit" of the molecular formula.**

Example:

The empirical formula of C_6H_6 would be: ______

1.3.2

Example: Empirical Formulas

What are the empirical formulas of the following molecules?

$C_6H_{12}O_6$: ________

$C_5H_{12}O$: ________

N_2O_4 : ________

Example: Empirical and Molecular Formula (Problem Type #1)

(Finding the Empirical and Molecular Formula Given Percent Compositions)

WIZE TIP

To find the empirical and molecular formulae given percent compositions, follow these steps:

Step 1 – Find the **mass %** composition of **each compound**

Step 2 – Assume 100 g sample and **calculate the grams of each element**

Step 3 – Calculate the **moles of each element**

Step 4 – **Divide** each number of moles by the **greatest common factor**

Step 5 – Write out the **empirical formula!**

Step 6 – If given the **molecular weight** of the molecule you can now find the **molecular formula**

- molar mass of empirical formula(x) = molar mass of molecule
- Multiply each subscript in the empirical formula by x to get the molecular formula!

Analysis of an unknown compound tells us that a sample has 39% C, 10% H, and 51% Oxygen. The molar mass of the compound was found to be 124g/mol. What is the empirical and molecular formula for this compound?

Step 1 – Find the mass % composition of each compound

Here we already have been given the percent composition by mass of each element:

C makes up ______%, H makes up ______%, and O makes up ______% of the sample by mass

Step 2 – Assume 100 g sample and calculate the grams of each element

If we assume a 100g sample, how many grams would be accounted for by C, H, and O?

C=______g H=______g O=______g

Step 3 – Calculate the moles of each element $n = \frac{m}{M}$

Step 4 – Divide each number of moles by the greatest common factor

Step 5 – Write out the empirical formula!

Now we know that for every _____ C, there are _____ H and _____ O

Based on this information, we can write the empirical formula: __________

Step 6 – If given the molecular weight of the molecule you can now find the molecular formula

- We are told the molar mass of the whole molecule is: 124g/mol
- If we find the molar mass of the empirical formula, we can see if we need to multiply it by a certain number to get the molar mass of the overall compound:

1.3.4

Example: Empirical And Molecular Formula (Problem Type #2)

(Finding the Empirical and Molecular Formulae Given Masses in Grams of Elements)

A sample of a compound contains 1.52g of N atoms and 3.47g of O atoms. The molar mass of the compound is between 90.0g and 95.0g. Determine the empirical and molecular formulas. Also, calculate the actual molar mass of this compound.

1.3.5

Practice: Molecular Formula from Empirical Formula

p-dichlorobenzene (mM = 147.00 g/mol) is a common disinfectant and deodorant used in industrial applications. It has an empirical formula of C_3H_2Cl. Determine its molecular formula.

Note: if your answer is C_3H_2Cl, enter it as: C3H2Cl.

Answer

Practice: Determine the Empirical and Molecular Formulae

A compound is found to contain 50.05% sulfur and 49.95% oxygen by weight. The molecular weight for this compound is 64.07 g/mol.

Part 1

What is the empirical formula for this compound?

A) SO

B) S_2O

C) SO_2

D) SO_3

Practice: Determine the Empirical and Molecular Formulae

A compound is found to contain 50.05% sulfur and 49.95% oxygen by weight. The molecular weight for this compound is 64.07 g/mol.

Part 2

What is the molecular formula for this compound?

A) S_2O_2

B) SO_4

C) S_2O

D) SO_2

1.4 Composition of Mixtures

1.4.1

Mass Percentage

Shows the amount that each element in a molecule contributes to the overall molecular mass.

WIZE TIP

To determine the mass percentage for a particular element, follow these steps:

1) Find the **atomic mass contribution of the element** in question

2) Find the **molar mass of the molecule**

3) **Solve for the mass percentage** of an element using the equation below (take the mass contribution of an element and divide by the total molar mass, then multiply by 100)

$$\boxed{\text{mass percentage of element} = \frac{\text{mass controbution of element}}{\text{total mass of molecule}} \times 100}$$

Example: What is the percentage of oxygen by mass in sodium hydroxide, NaOH?

Therefore, oxygen contributes to ______% of the overall mass of NaOH!

1.4.2

Example: Mass Percentages of Elements in A Molecule

Sodium Bicarbonate ($NaHCO_3$) is a very common chemical frequently used in the food industry. Find the mass percentages of each element in sodium bicarbonate.

1.4.3

Practice: Finding the Mass Percentage of Oxygen in Ethanol

What is the mass percentage of oxygen in ethanol (C_2H_5OH)?

- A) 31.1 %
- B) 34.7 %
- C) 35.5 %
- D) 41.1 %

1.5 Hydrate Practice

1.5.1

A 7.21 g sample of a $LiClO_4 \cdot xH2O$ was heated to drive off the water. The mass was reduced to 4.78 g. What is the formula of the hydrate?

- $LiClO_4 \cdot H_2O$
- $LiClO_4 \cdot 3H_2O$
- $LiClO_4 \cdot 4H_2O$
- $LiClO_4 \cdot 7H_2O$

1.5.2

A 15.67 g sample of a hydrate of magnesium carbonate ($MgCO_3$) was heated, without decomposing the carbonate, to drive off the water. The mass was reduced to 7.58 g. What is the formula of the hydrate?

- $MgCO_3 \cdot H_2O$
- $MgCO_3 \cdot 2H_2O$
- $MgCO_3 \cdot 4H_2O$
- $MgCO_3 \cdot 5H_2O$

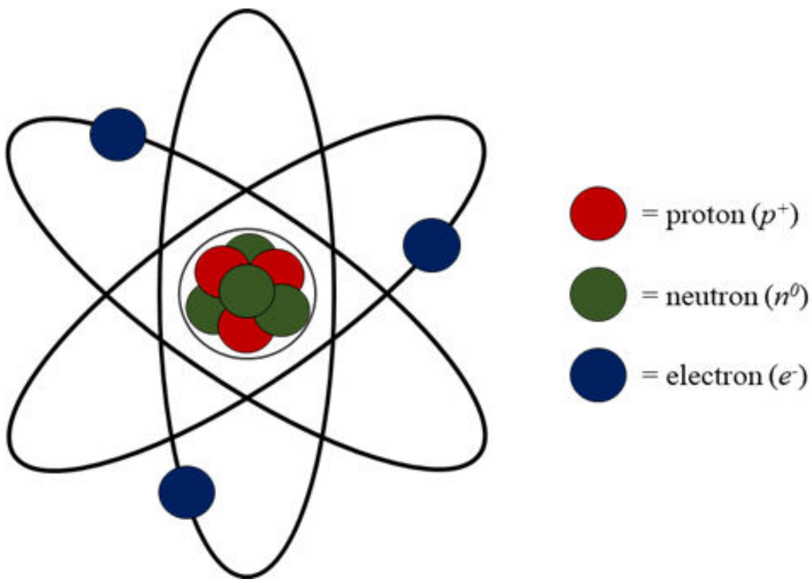
1.6 Atomic Structure

1.6.1

The Atom and its Subatomic Particles

Atoms are the building blocks of chemistry (and life!). They make up everything from the screen of your computer, to components in your eyes! They are all around us. Since atoms are such a key concept in chemistry, let's take a closer look at their structure.

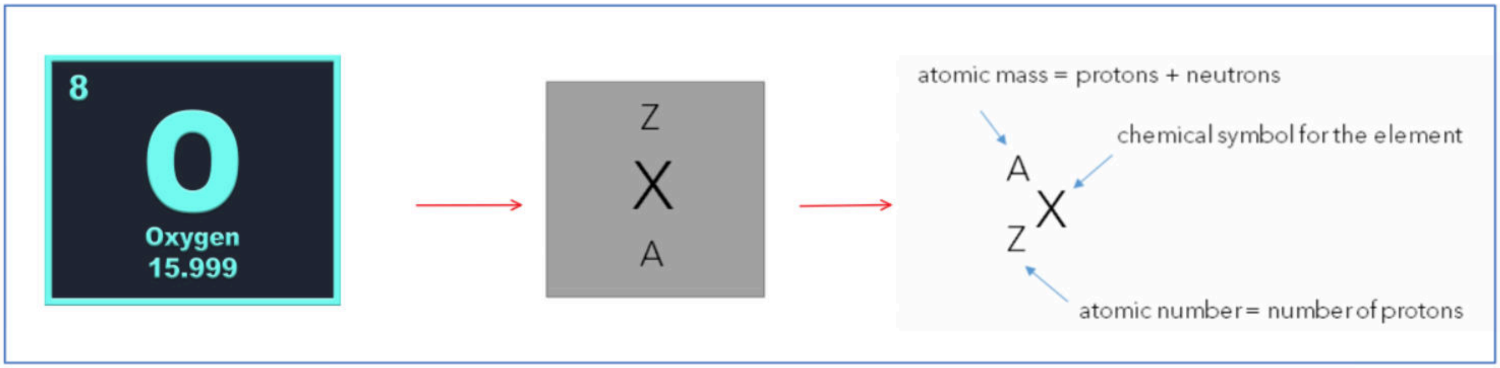
Atomic Structure



Subatomic Particles

Particle	Mass (g)	Mass (amu or $g{\cdot}mol^{-1}$)	Relative Charge
Proton	1.673×10^{-24}	1.007	+1
Neutron	1.675×10^{-24}	1.009	0
Electron	9.109×10^{-28}	5.485×10^{-4}	-1

Chemical Symbols and Notation



1.6.2

Example: Nuclear Notation

Ex1) ${}^{12}_{6}C$

# of protons: _____

# of neutrons: _____

# of electrons: _____

WIZE CONCEPT

Atomic Mass = # of protons + # of neutrons
# of neutrons = Atomic mass - # of protons

Ex2) ${}^{19}_{9}F^{-}$

# of protons: _____

# of neutrons: ______

# of electrons: ______

WIZE TIP

When **electrons are added** to an element we call it an **anion**, and it would have a **negative charge**.
When **electrons are removed** from an element we call it a **cation**, and it would have a **positive charge**

1.6.3

Practice: Nuclear Notation

How many protons, neutrons and electrons are there in $^{75}As^{3-}$?

Number of protons

Number of neutrons

Number of electrons

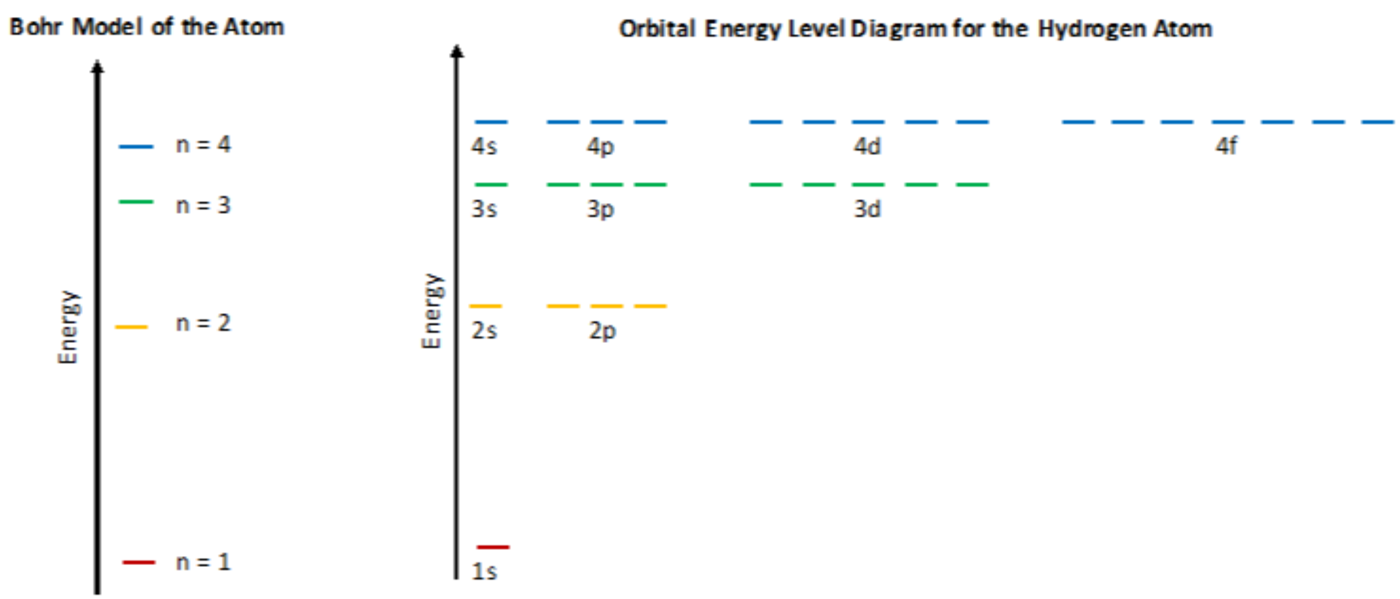
1.7 Orbital Filling Diagrams

1.7.1

Relative Energies of Atomic Orbitals

One-Electron Species

- For a one-electron species like the hydrogen atom, there is no electron-electron repulsion. Therefore, all the subshells of the same n are **degenerate.**
- Orbitals at the **same energy level** are called **degenerate.**

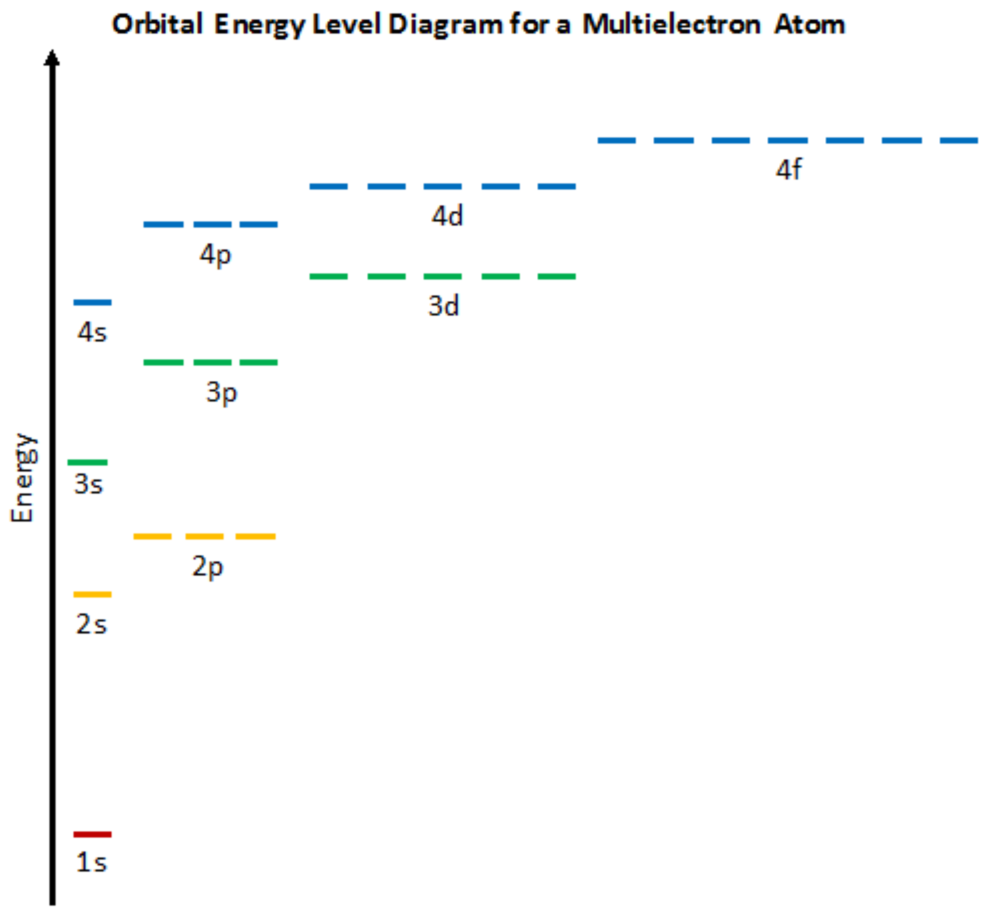


WIZE TIP

It will be helpful to understand the difference between a **shell**, a **subshell**, and an **orbital.**
We will label these in the diagram above!

Multi-Electron Species

- With multi-electron species, there are interactions between electrons (electron-electron repulsion)
- The energy level diagram of the orbitals looks like this:



WIZE CONCEPT

- S subshells hold a maximum of _____ electrons
- P subshells hold a maximum of _____ electrons
- D subshells hold a maximum of _____ electrons
- F subshells hold a maximum of _____ electrons

$$\boxed{\text{The total number of orbitals in a given shell is given by : } n^2}$$

Example:

If n=2, what is the total number of orbitals in that shell?

- ________, circle the _____ orbitals that have n=2 above

What if we wanted to know the maximum number of electrons we could have if n=2?

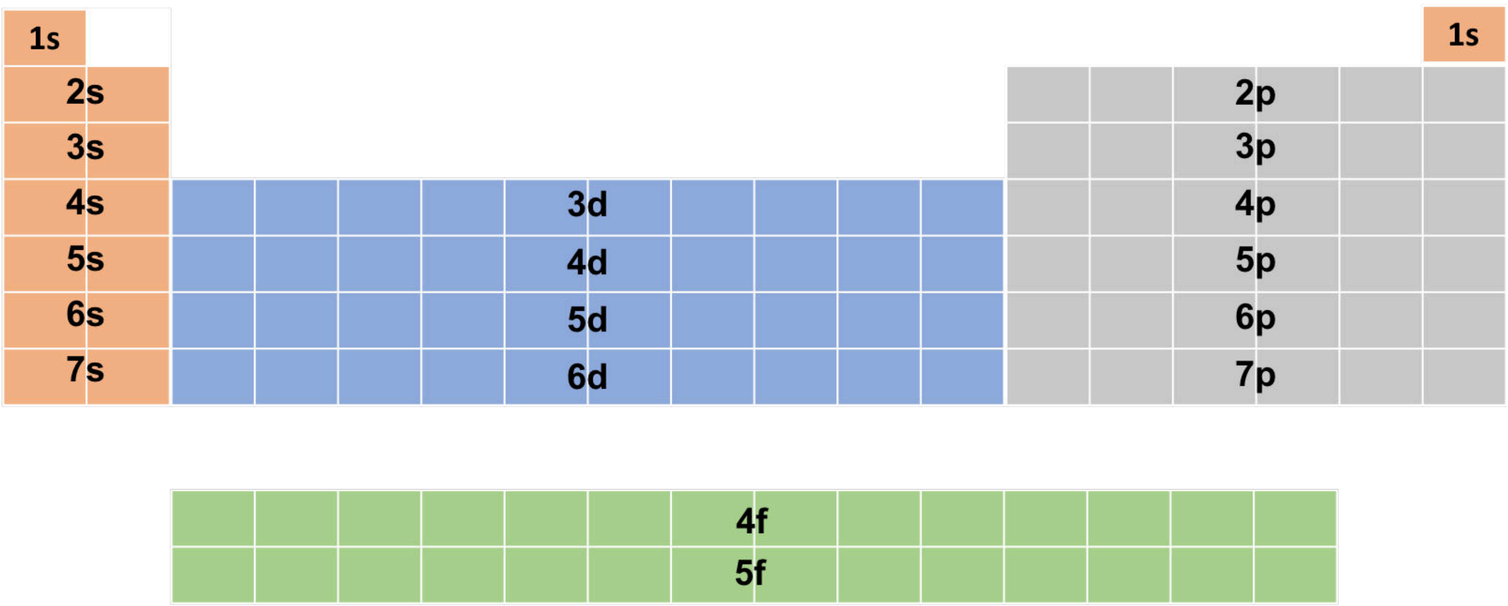
- We already found that total number of orbitals for this shell (_____)
- What is the maximum number of electrons we could have in each orbital? _____
- Therefore the maximum number of electrons possible when n=2 is _____

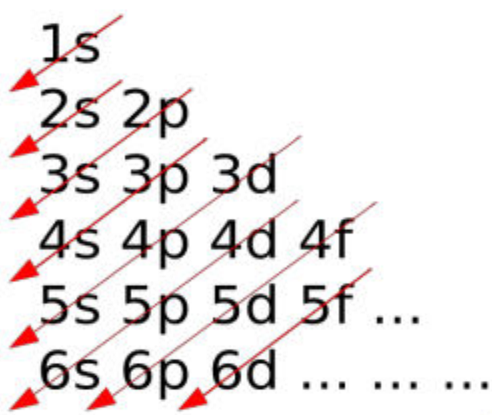
1.7.2

Rules for Orbital Filling

1) Aufbau Principle

Electrons will always occupy the **lowest available energy level first.**





WATCH OUT!

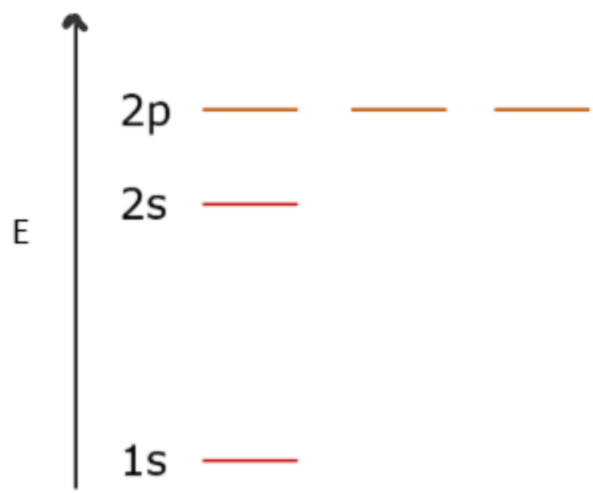
Memorize the order of orbital filling (or just familiarize yourself enough with the periodic table to know the order!): 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p...

2) Hund’s Rule

Due to electron-electron repulsion, electrons will **fill orbitals of the same energy singly before pairing up**

Electrons don't want to be next to each other unless they have to be!

Example: Fill out the following orbital diagram for C



3) Pauli Exclusion Principle

No two electrons in an atom will have the same set of **4 quantum numbers.**

Quantum Number	Values	Interpretation
m_s (the spin quantum number)	-1/2 or +1/2	An electron behaves like a magnet that has one of two possible orientations, aligned either with the magnetic field or against it

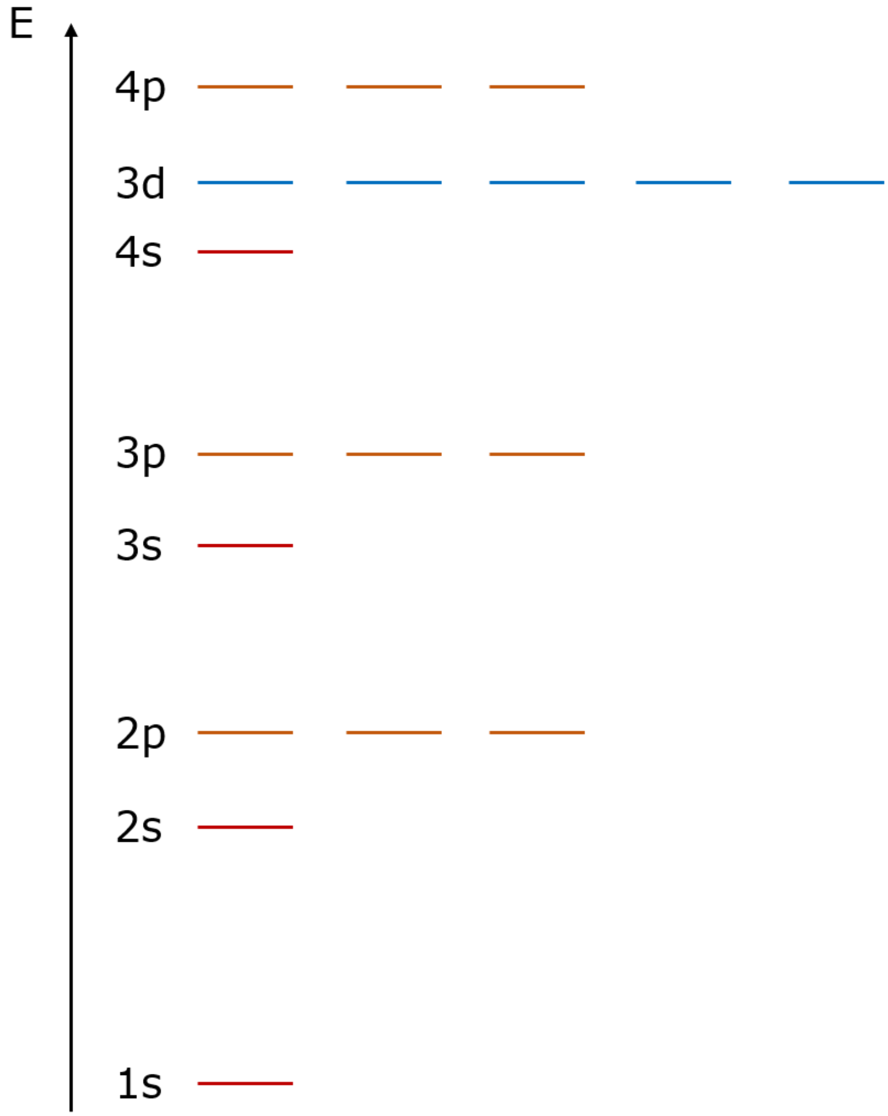
1.7.3

Example: Orbital Filling Diagrams

Draw the orbital diagram for oxygen.

Example: Orbital Filling Diagrams

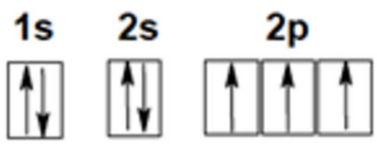
Complete the atomic orbital diagram below for a neutral calcium atom.



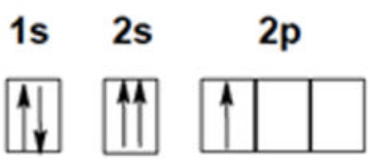
Practice: Orbital Filling Diagrams

Which of the orbital diagrams gives the correct electron configuration for an atom of boron, in the ground state?

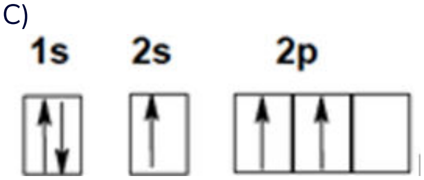
A)



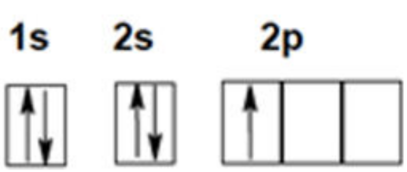
B)



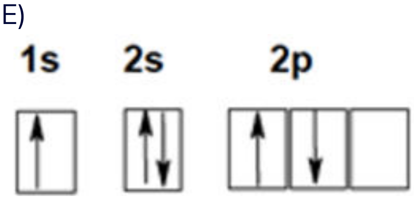
C)



D)



E)



Practice: Number of Orbitals

How many different atomic orbitals exist where n = 3?

- 3
- 6
- 9
- 12, which
- 18

1.8 Electronic Configuration for Atoms

1.8.1

Electron Configurations of Atoms

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

- lable blocks
- describe the n number for each block

Simple Electron Configurations

Li:

C:

Ne:

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

Shorthand Notations

Write the name of the **previous noble gas and then fill in the rest of the electron configuration.**

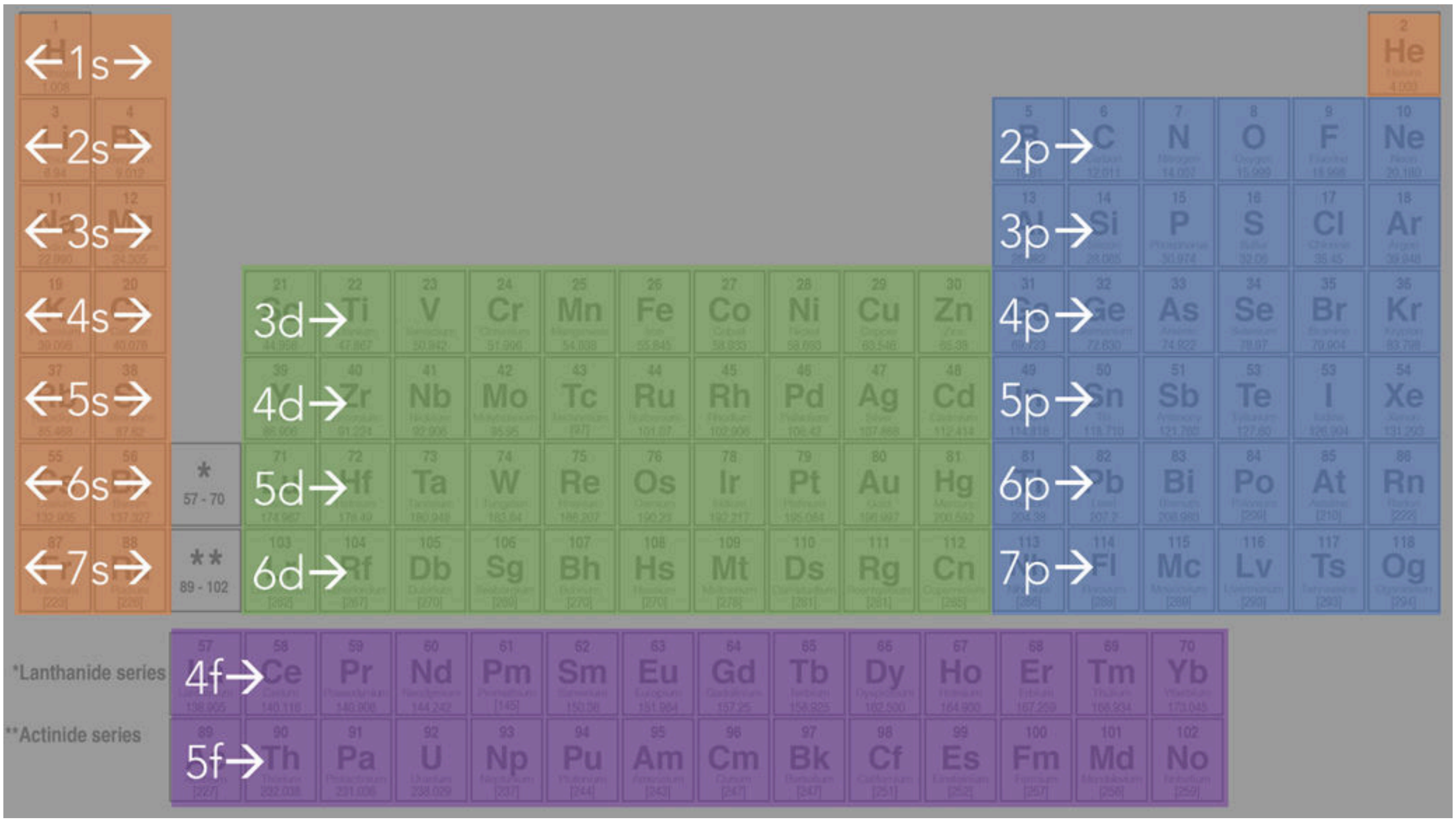
O:

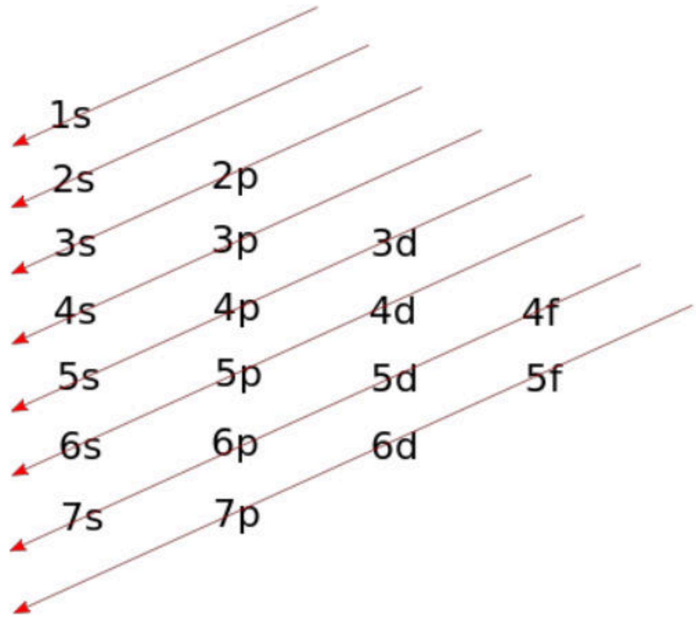
Fe:

Ge:

1.8.2

Electron Configurations Cheatsheet




1s
2s
2p
3s
3p
3d
4s
4p
4d
4f
5s
5p
5d
5f
6s
6p
6d
7s
7p

1.8.3

Example: Writing Out Electron Configurations

Part 1:

Write the full electronic configurations for the following elements:

P:

Se:

Part 2:

Write the short electron configurations for the following elements:

Zn:

Zr:

The Periodic Table of the Elements

Group→ / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

1.8.4

Practice: Identify the Neutral Element

What neutral element is represented in the example below?

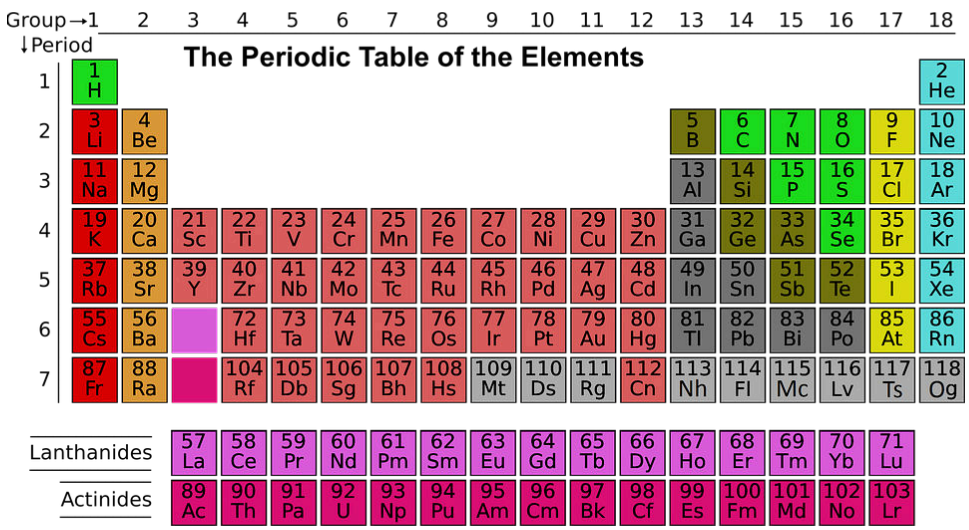
$1s^2 2s^2 2p^6 3s^2 3p^5$ became $[Ne]3s^2 3p^5$

- Al
- Ne
- P
- Cl

1.8.5 Practice

Practice: Impossible Electron Configuration

Which of the following electronic configurations is not possible?



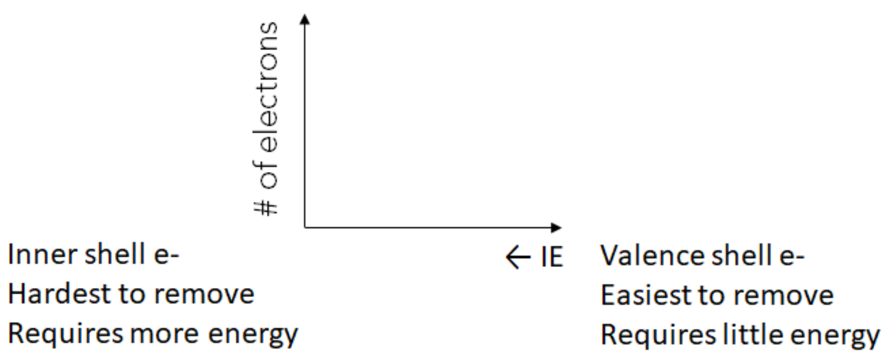
- $[Ar]4s^2 4d^6$
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
- $[Kr]5s^1$
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

1.9 Photoelectron Spectroscopy

1.9.1

Photoelectron Spectroscopy

- Photoelectron spectrophotometers use high-energy radiation (UV or X-rays) to eject electrons from an atom
- Since electrons within an atom are in different energy levels, different electrons can require different amounts of energy to eject
- Electrons that are CLOSER to the nucleus are HARDER to remove (greater attraction between protons and electrons)
- VALENCE electrons (outermost electrons) are the EASIEST to remove
- The photoelectron spectrophotometer removes electrons from multiple atoms, so the electrons from all levels will are observed and shown on the resulting graph --> we can determine the electron configuration of atoms!



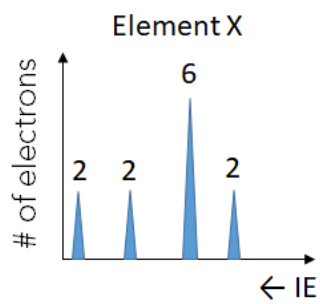
Draw the photoelectron spectrum for O and use it to explain the common bonding pattern for the oxygen atom.

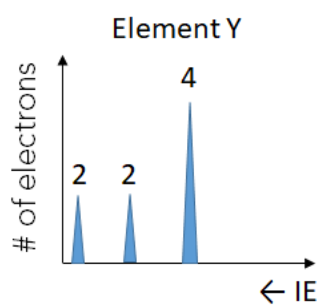
Compare this to the nitrogen atom?

1.9.2

Ionic Compounds and Photoelectron Spectroscopy

The PES for a compound made up of X and Y elements in shown below. Which of the following is the most likely chemical formula of this compound?





1.10 Effective Nuclear Charge (Zeff)

1.10.1

Effective Nuclear Charge (Zeff)

To understand effective nuclear charge (Zeff), let's consider a concert:

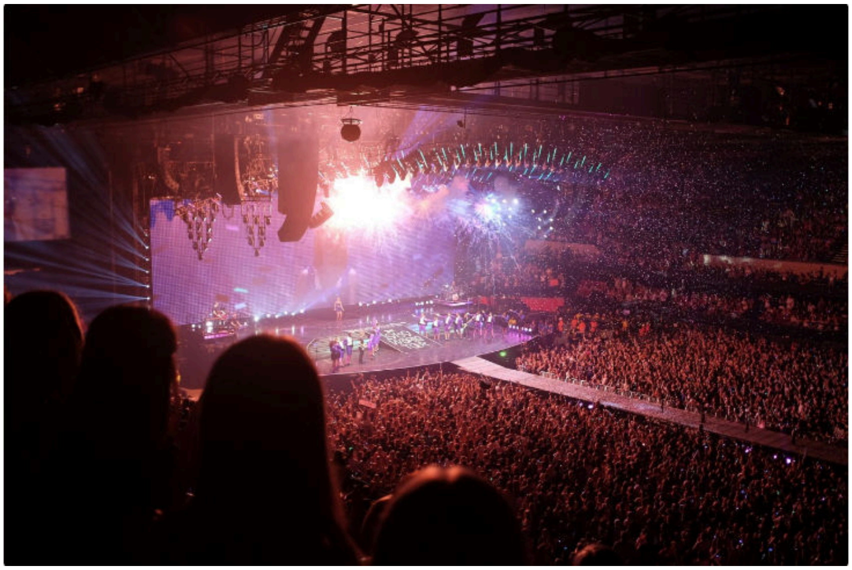


Photo by The Come Up Show / CC BY

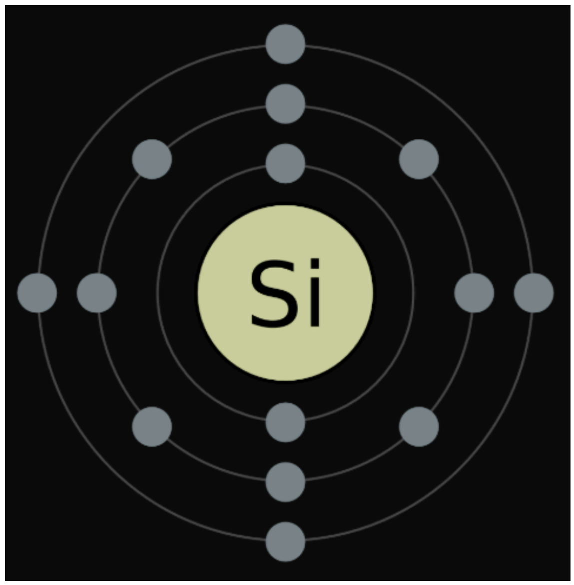
- People closer to the stage at a concert are going to be more into it. The music is louder and they really get a close connection with the artist. People in the highest rows in the stadium get less of that connection with the artist performing.
- This helps to explains the concept of nuclear shielding. **Inner electrons** (or people sitting in the front rows in our example) **shield the outer electrons** (or people sitting in the highest rows) from the **attractive force of the nucleus** (or artist in our example)

Effective Nuclear Charge

This is the **nuclear charge that is "felt"** by a valence electron.

Core Electrons Vs Valence Electrons

Label the **core** and **valence electrons** in the diagram below:



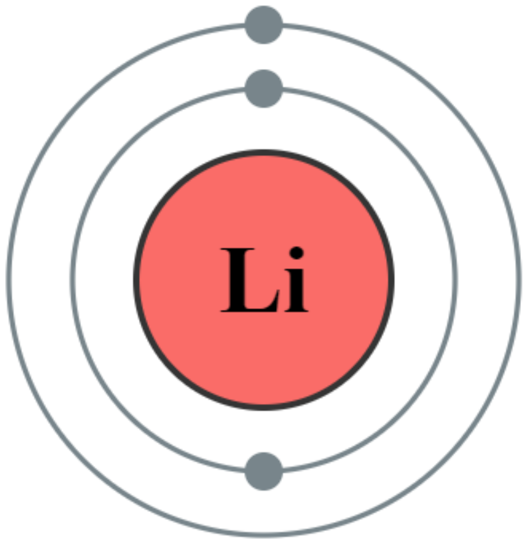
- Valence electrons are attracted to the positively charged nucleus BUT valence electrons are repelled by the core electrons
- Note that electrons in the same shell "feel" the same attraction to the nucleus (since they are they same distance from the nucleus, just like how the people in the same row would feel the same connection to the artist)

$$\boxed{Z_{eff} = Z - S}$$

Zeff *is the* ***effective nuclear charge***
Z *is the* ***atomic number*** *(# of protons)*
S *is the # of* ***shielding electrons***

Are the shielding electrons core electrons or valence electrons? ________

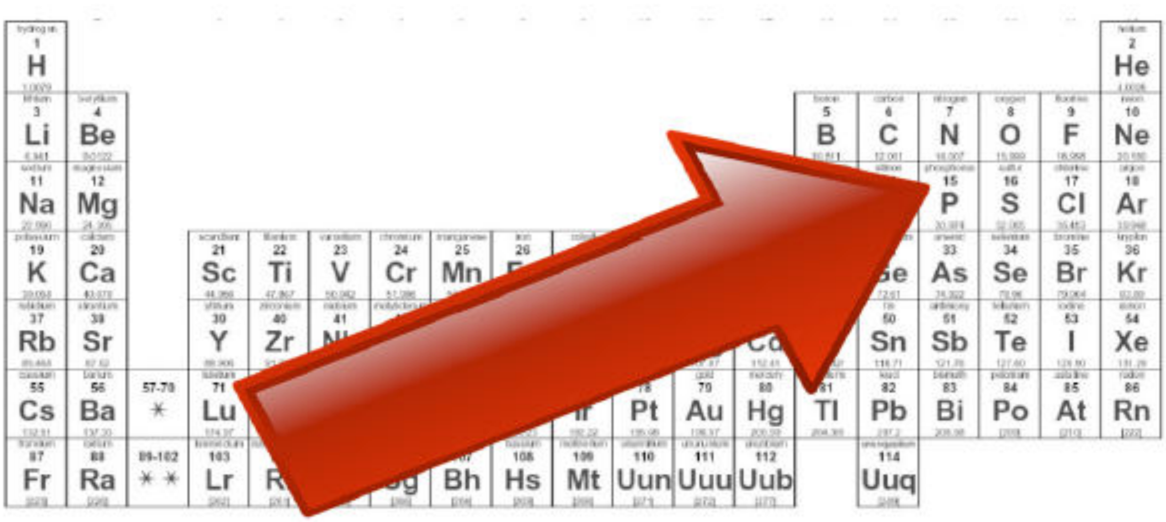
What would be the effective nuclear charge for Lithium?



The Periodic Trend

- As we **move to the right** across the periodic table, the # of core electrons stays the same but the **# of protons increases**. Therefore, **Z_{eff} increases.**
  - With protons, can pull the electrons in closer → Zeff
  - Effects of shielding are less with more protons

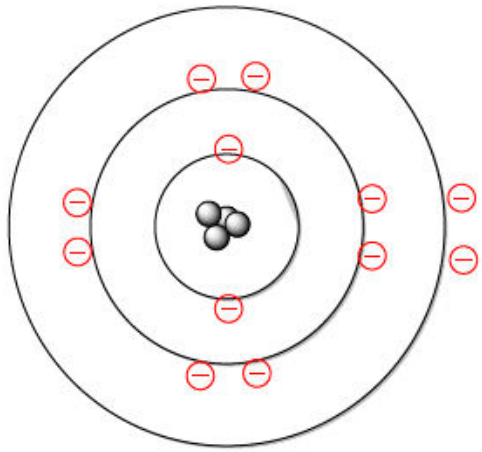
- As we **move down a group** of the periodic table, the **# of core shells increases** and the valence electrons get further from the nucleus. This increased distance from the nucleus leads to a **smaller Z_{eff}.**
  - a group → shells
  - shells → shielding
  - shielding → Zeff



1.10.2

Example: Estimating Zeff

What is the Z_{eff} of an electron in the n=3 shell of magnesium (Mg)?



1.10.3 **Practice**

Practice: Estimating Zeff

Calculate the approximate effective nuclear charge of Se.

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

- 6
- 8
- 28
- 34

1.11 Atomic Radius Trend

1.11.1

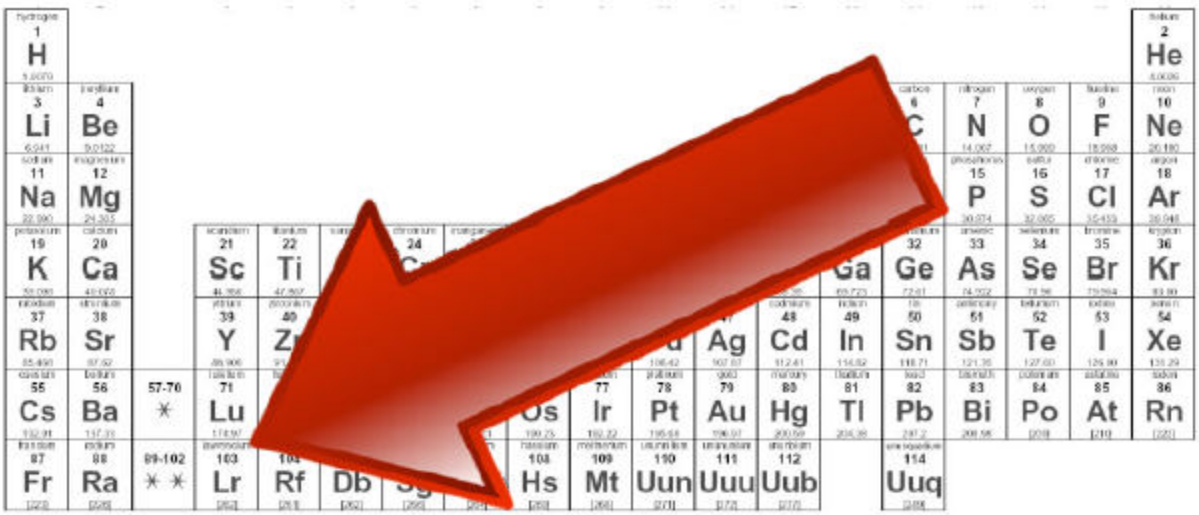
Atomic Radius

The atomic radius is the estimated **radius of an atom** (from the nucleus to the outermost valence electrons)

The Periodic Trend

- As we **move to the right** on the periodic table, the **Z_{eff} increases** and this “pulls” the electrons closer to
  the nucleus which **decreases the radii.**

- As we **move down a group** in the periodic table, the number of **electron shells increases** which makes the
  atom radii larger.
  - group → shells
  - shells → shielding and Zeff
  - With a Zeff → pull on outer electrons, leading to atomic radii



Element with the **largest atomic radii** in the periodic table: ______

Element with the **smallest atomic radii** in the periodic table: ______

How do ions change this?

- **Anions:** __________ due to increased electron-electron repulsion
- **Cations:** __________ due to decreased electron-electron repulsion
  - How does this relate to Zeff?

In general:

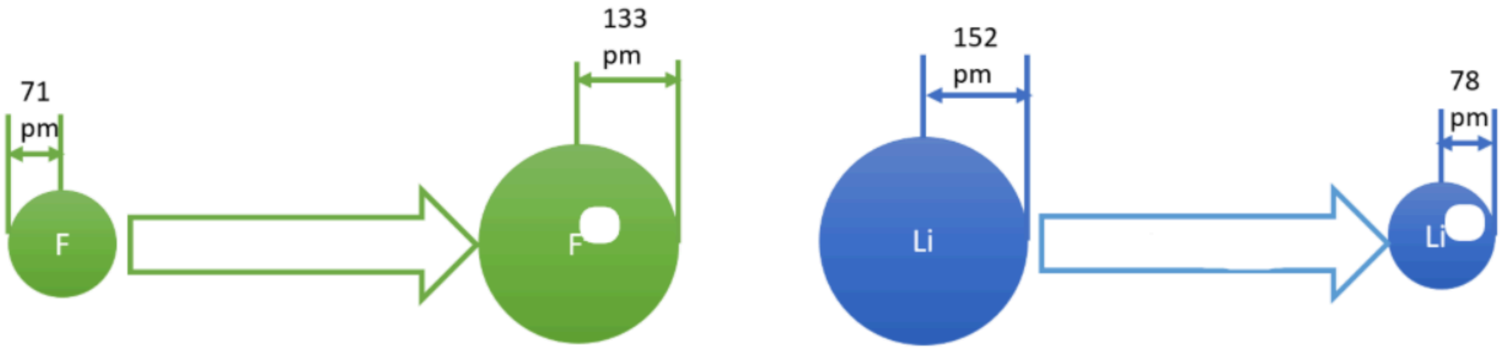
1.11.2

Ionic Radius: Three Scenarios

You could be asked to rank different atoms or ions according to the sizes of their ionic radii using the trends we discussed.

Here are three common scenarios you may encounter:

1. Same element different charge:



2. Different element same charge:

- Identical trend to atoms
  Example: Li < Na so Li^+ < Na^+

3. Different element different charge

- Can only be assessed for **isoelectronic species** (same # of electrons)
- Compare the **proton to electron ratio.**
  - **More protons than electrons**, means stronger pull, **smaller radius.**
  - **Less protons than electrons**, means weaker pull, **larger radius.**

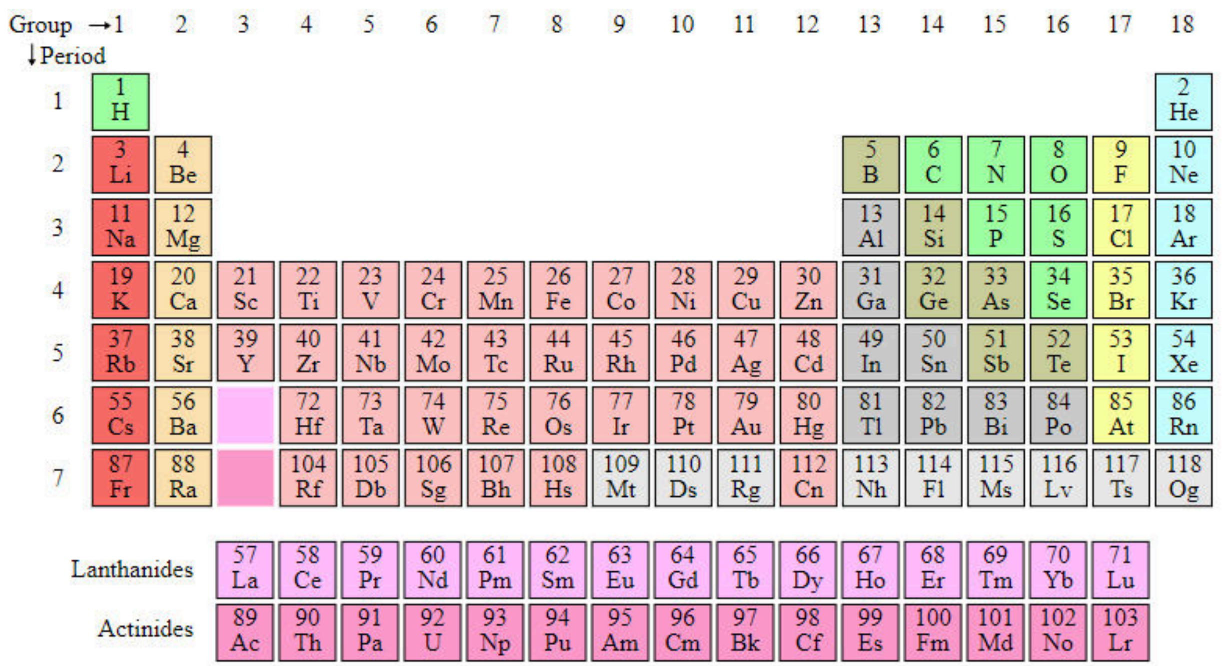
  Example: Compare the radii for the following: O^{2-} F^- Ne Na^+ Mg^{2+}

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	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			89 Ac	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	

1.11.3

Example: Ranking Size of Atoms

Rank the following atoms in order of increasing atomic radius: Se, Cs, Br, Ga, F, As.

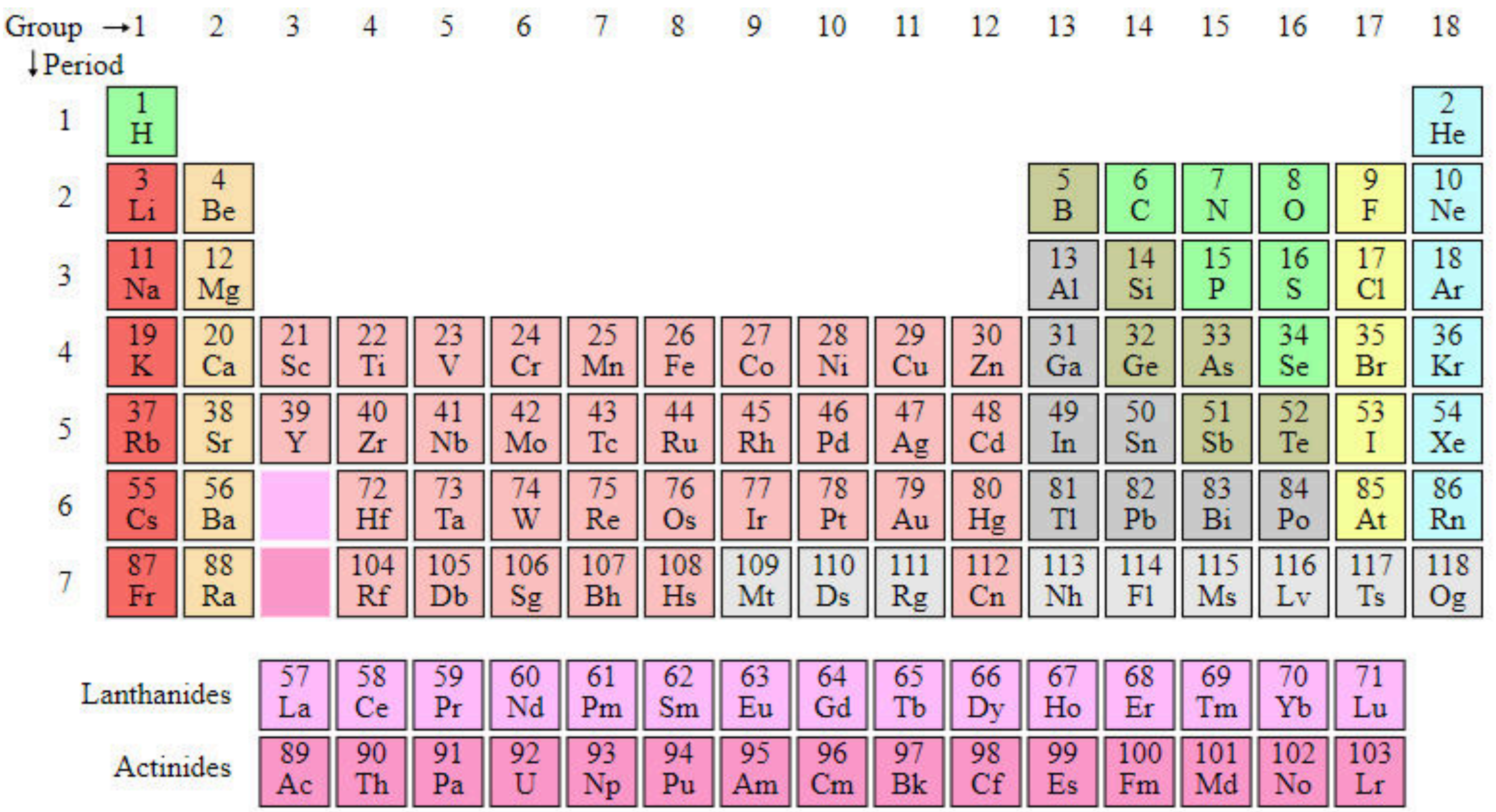


Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

Example: Ranking Size of Ions

Rank the following species in order of decreasing size. F^{-}, N^{3-}, Al^{3+}



WIZE TIP

When comparing atoms that are **isoelectronic** (i.e. they have the same number of electrons), remember **more protons in the nucleus** means a stronger the pull of electrons toward the + charged nucleus, so the **smaller the atomic radius**

1.11.5

Practice: Determining the Largest Atom

Select the atom below which has the largest atomic radius?

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

- S
- I
- O
- In
- Rb

1.12 Ionization Energy

1.12.1

Ionization Energy (IE)

The ionization energy is the amount of **energy required to remove the outermost electron** from an atom or ion in the gas phase

- If removing an electron do we get an anion or a cation? ________

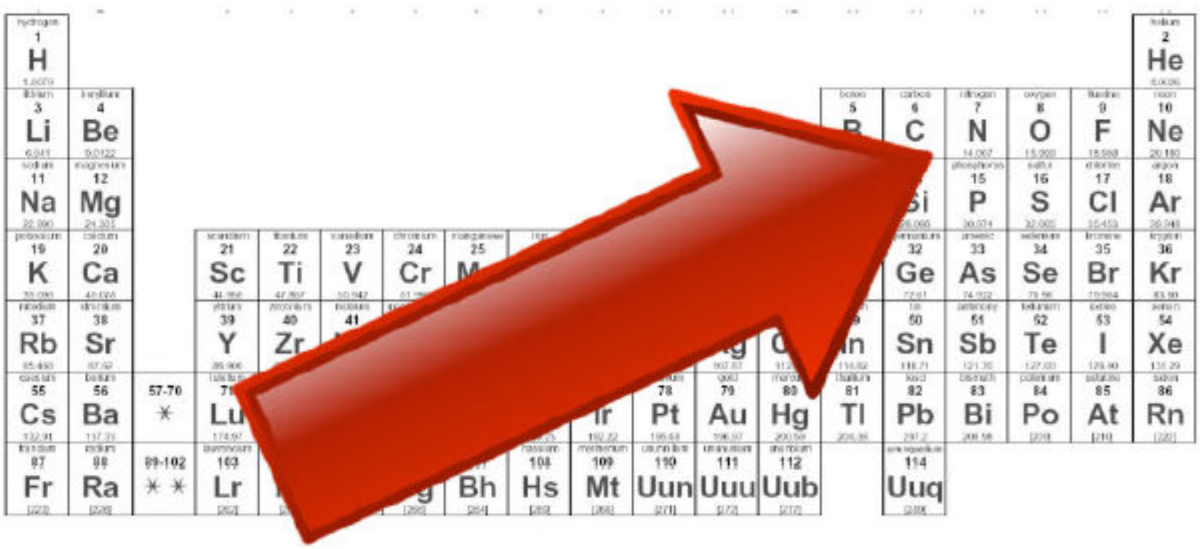
$$\mathbf{E} \rightarrow \mathbf{E}^{+} + \mathbf{e}^{-} \qquad \Delta \mathrm{E} = \text{Ionization Energy}$$

- **↑IE** means that the atom/ion requires **more energy** for a valence **electron to be removed** (has a tight pull on electron)

The Periodic Trend

- As we **move to the right** across a period, **Z_{eff} increases** and the electrons are held more tightly. Therefore, it takes more energy to remove an electron and the **IE increases** as we move to the right.

- As we **move down a group** and the valence electrons are further away from the nucleus, they are held more weakly (**lower Zeff)** and **IE decreases** (it becomes easier to remove the outermost electron)


H He
Li Be C N O F Ne
Na Mg P S Cl Ar
K Ca Sc Ti V Cr Ge As Se Br Kr
Rb Sr Y Sn Sb Te I Xe
Cs Ba 57-70 Lu Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra 89-102 Lr Bh Hs Mt Uun Uuu Uub Uuq

Comparing 1st, 2nd, and 3rd Ionization Energies

How does the first IE compare to the second IE?

$$First\ Ionization\ Energy\ (IE_1):\quad A\ (g)\ \rightarrow\ A^+\ (g)\ +\ e^-$$

$$Second\ Ionization\ Energy\ (IE_2):\quad A^+\ (g)\ \rightarrow\ A^{2+}\ (g)\ +\ e^-$$

$$Third\ Ionization\ Energy\ (IE_3):\quad A^{2+}\ (g)\ \rightarrow\ A^{3+}\ (g)\ +\ e^-$$

1st ionization is always smaller than the 2nd because removing a negatively charged electron from a cation is more difficult.

Example:

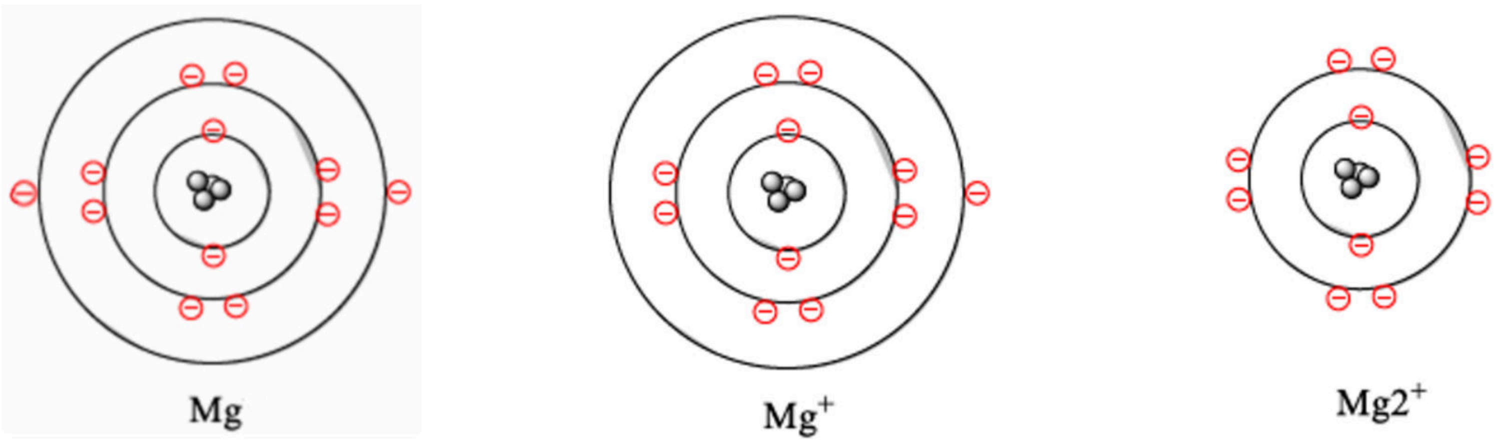
$$Ca\ (g)\ \rightarrow\ Ca^+\ (g)\ +\ e^-\quad IE_1\ =\ 590\ \frac{kJ}{mol}$$

$$Ca^+\ (g)\ \rightarrow\ Ca^{2+}\ (g)\ +\ e^-\quad IE_2\ =\ 1145\ \frac{kJ}{mol}$$

WIZE TIP

IE_1 < IE_2 < IE_3 < IE_4 etc

How does the IE_2 and IE_3 for magnesium compare?



A **core electron is much more difficult to remove than a valence electron**, for example: the 3rd ionization energy of magnesium is much, much, higher than the 2nd.

> **WIZE TIP**
>
> Although the **IE increases each time another electron is removed**, **the increase is not linear** and is usually **related to the electron configuration.**

Ionizing Mg^{2+} would be very difficult!

$$Mg \rightarrow \ Mg^{+} \ + \ e^{-} \quad IE_1 \ = 738 \ \frac{kJ}{mol}$$

$$Mg^{+} \rightarrow \ Mg^{2+} \ + \ e^{-} \quad IE_2 \ = 1451 \ \frac{kJ}{mol}$$

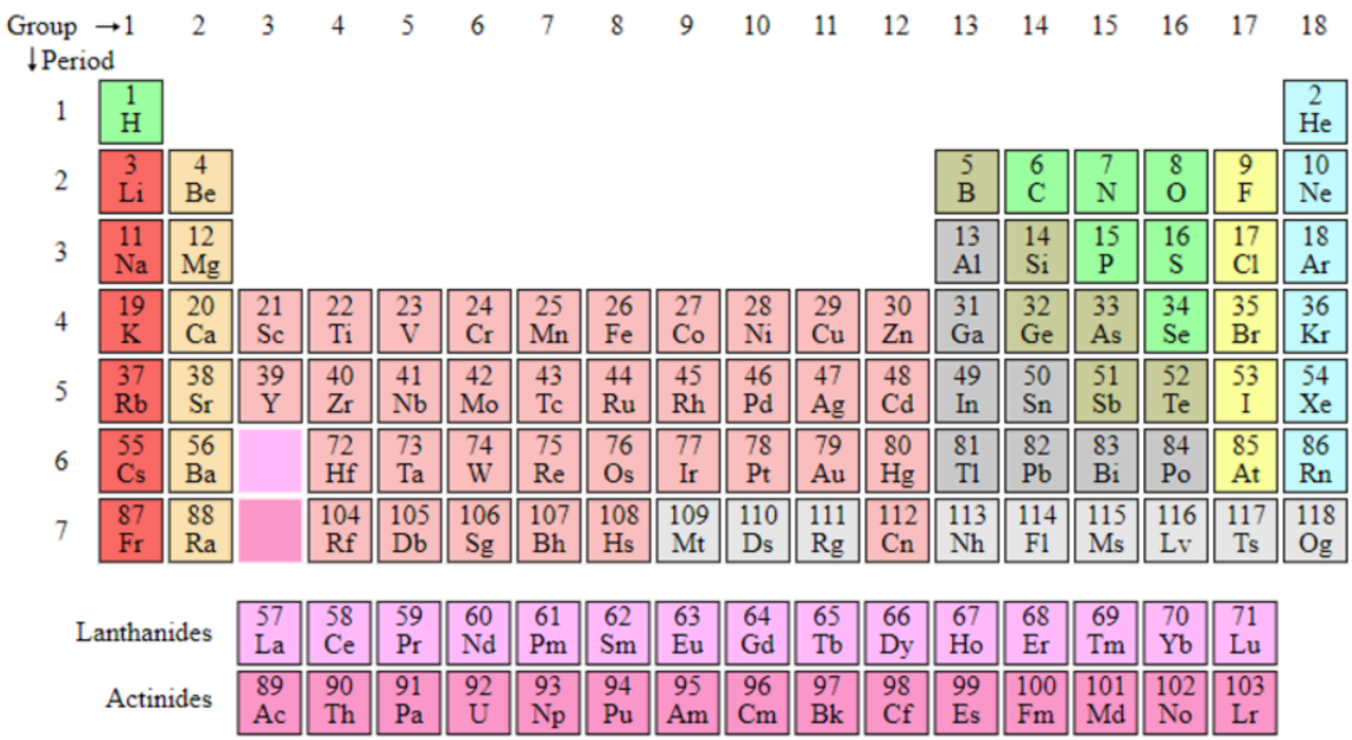
$$Mg^{2+} \rightarrow \ Mg^{3+} \ + \ e^{-} \quad IE_3 \ = 7733 \ \frac{kJ}{mol}$$

Needing so much energy to remove the 3rd electron indicates that when 2 were removed it was stable/unstable: __________!

1.12.3

Practice: Highest 1st Ionization Energy

Which of the following species has the highest 1st ionization energy?



- S
- F
- B
- N
- P

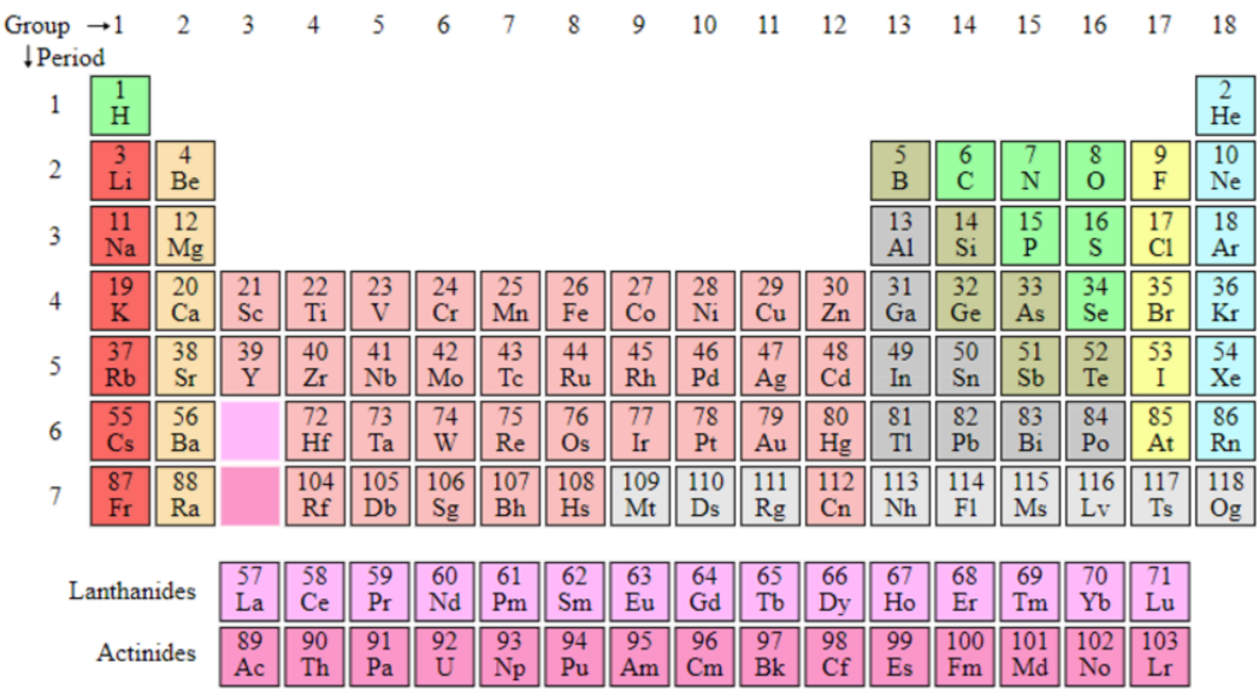
View Solutions on Wizeprep.com

Solutions to these questions, as well as step-by-step breakdowns of the answers at:

1.12.4

Practice: Trends in Ionization Energy

Rank the following atoms in order of increasing first ionization energy (1 = smallest; 6 = largest): Fe, O, Ba, Si, V, Zr



Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
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Actinides	89 Ac	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

Fe	
O	
Ba	
Si	

V	
Zr	

Practice: Lowest Ionization Energy

Select the element with the lowest 3rd ionization energy.

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

- **A)** Sc
- **B)** Mg
- **C)** Cl
- **D)** I
- **E)** C

1.13 Electron Affinity Trend

1.13.1

Electron Affinity (EA)

Electron affinity is the **amount of energy involved with adding an electron** to an atom or ion.

WIZE TIP

Don't confuse ionization energy (IE) with electron affinity (EA)!

- **Ionization energy** is when we **remove** the outermost electron
- **Electron affinity** is when we **add** an electron

Memory tip: Electron affinity has an "**a**" in it for adding an electron

When adding an electron, do we get a cation or an anion? ________!

$$\mathbf{E} + \mathbf{e}^{-} \rightarrow \mathbf{E}^{-} \qquad \Delta\mathbf{E} = \textbf{Electron Affinity}$$

- **If EA is negative** → energy is released/absorbed: ____________
    - If this is the case, would the atom go to a higher/lower E state? __________
    - This means the atom has become more/less ________ stable as a result of adding an electron
    - Did this atom like or dislike having an electron added to it? _______

- **If EA is positive** → energy is released/absorbed: ____________
    - If this is the case, would the atom go to a higher/lower E state? __________
    - This means the atom has become more/less stable________ stable as a result of adding an electron
    - Did this atom like or dislike having an electron added to it? __________

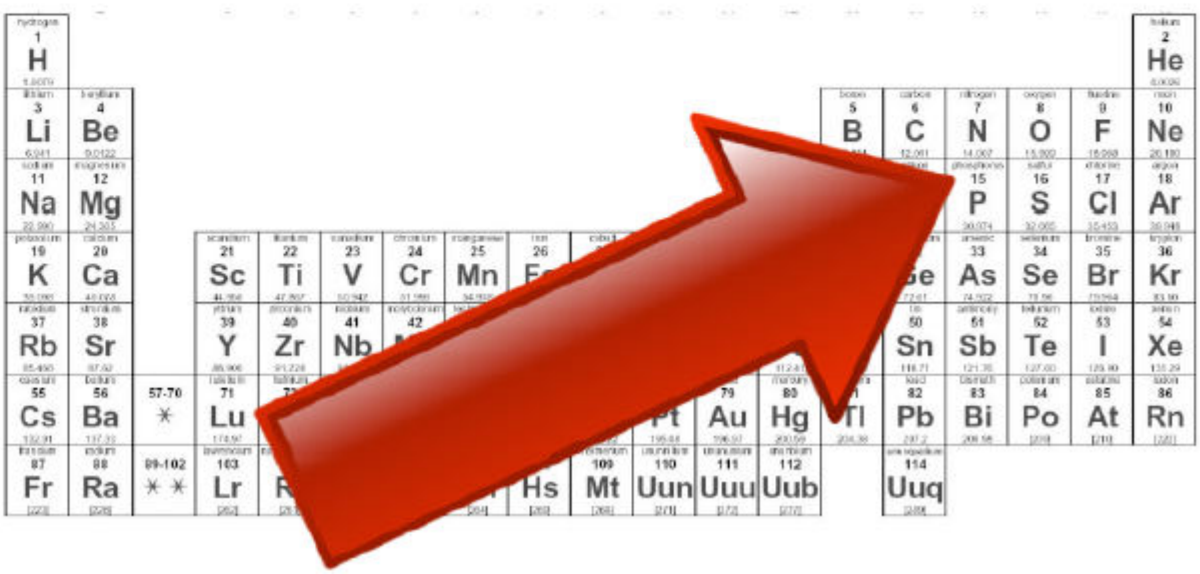
WATCH OUT!

Use a **negative EA as a reference point**, unless the question states otherwise.

- This means that unless specified otherwise, when asked which element would have the largest electron affinity, assume they mean largest NEGATIVE electron affinity.
    - This would mean that the element really liked getting an electron and lost energy, resulting in a more stable form.

The Periodic Trend

In general, as Z_{eff} increases, the **incoming electron** experiences a greater electrostatic attraction and **stabilization** leading to a greater (negative) electron affinity



What about noble gases?

Do you think a noble gas would have a positive or negative electron affinity? ____________

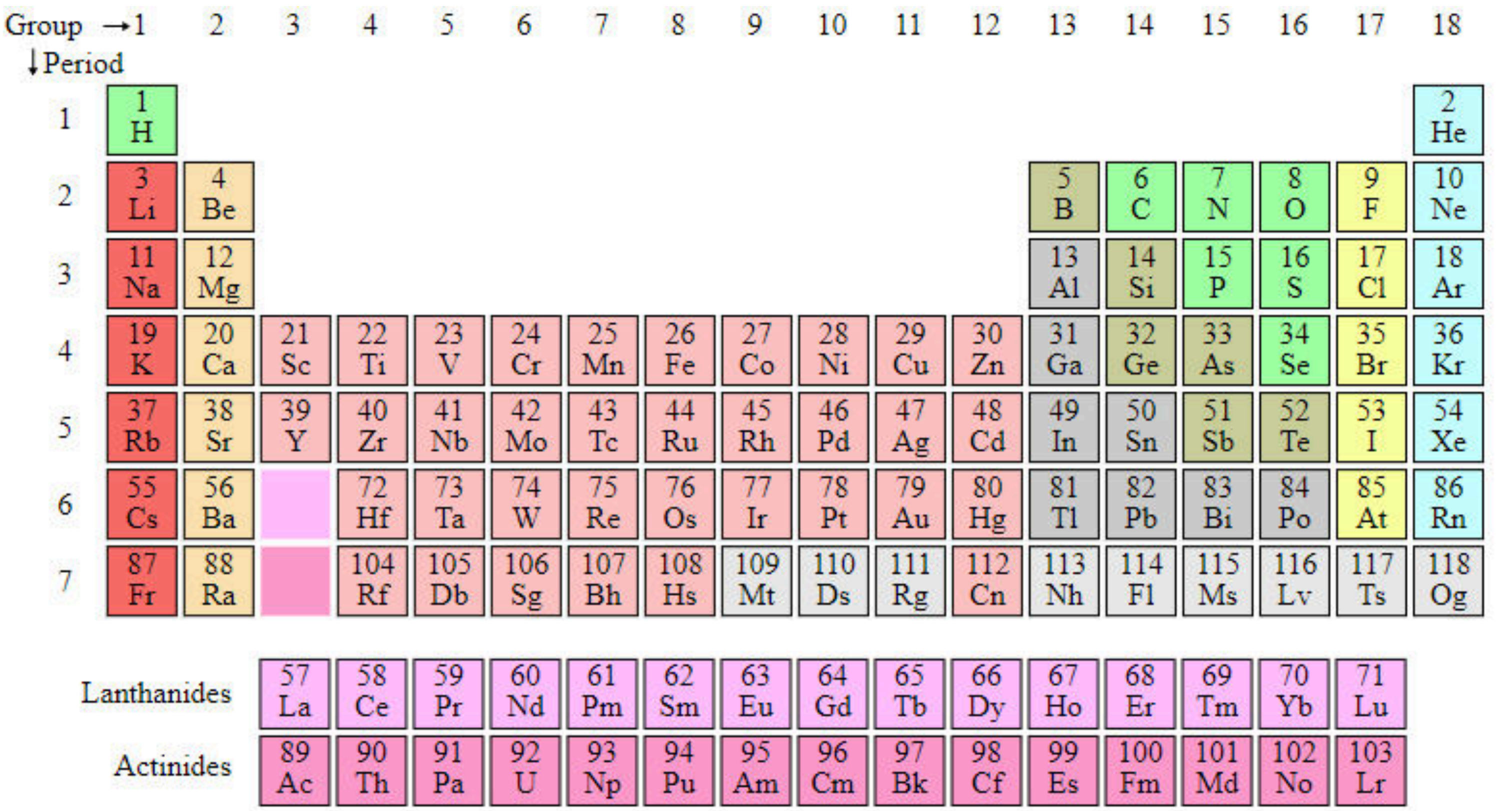
- Are noble gases stable? ________
- Would a noble gas want another electron added? ________
- If you tried to add an electron to a noble gas would that require energy or release energy?

  ____

1.13.2

Example: Increasing Electron Affinity

Rank the following atoms in increasing absolute energy of their first electron affinity: Br, Ag, Ba, Mo, Sb



1.13.3

Practice: Positive Electron Affinity

For most elements, the first electron affinity is negative; however, a few elements have a positive first electron affinity. Which atom has the most positive first electron affinity?

- F
- B
- Ne
- C
- O
- Na

1.13.4

Practice: Highest Electron Affinity

Which of the following has the highest electron affinity?

- N
- N+
- N-
- Not enough information is provided to determine this.

1.14 Electronegativity

1.14.1

Electronegativity (EN)

Electronegativity is a **measure of electron pull.** In a bond, the **more electronegative** element will **pull electrons closer towards itself**

Example: Draw the bond between H and F and show the difference in electronegativity.

Group → / ↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 1.87	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo
* Lanthanoids			La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27	
** Actinoids			Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3	

Electronegativity Values

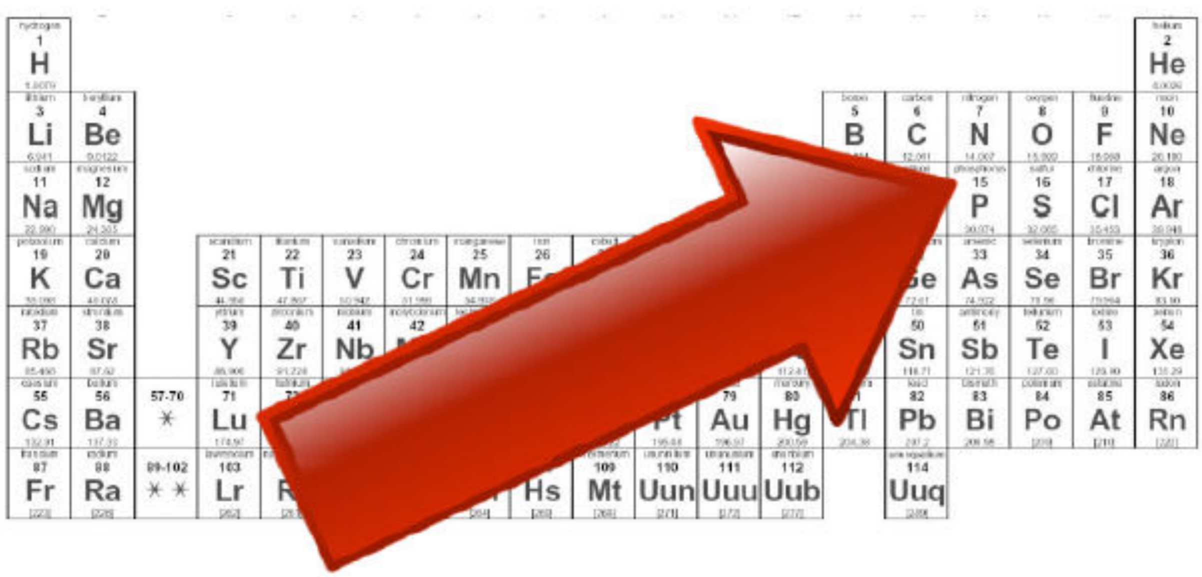
WIZE TIP

General order of EN may come in handy:

F > O > N > Cl > Br > I > S > C ~ H

The Periodic Trend

Electronegativity increases going up and to the right in the periodic table.

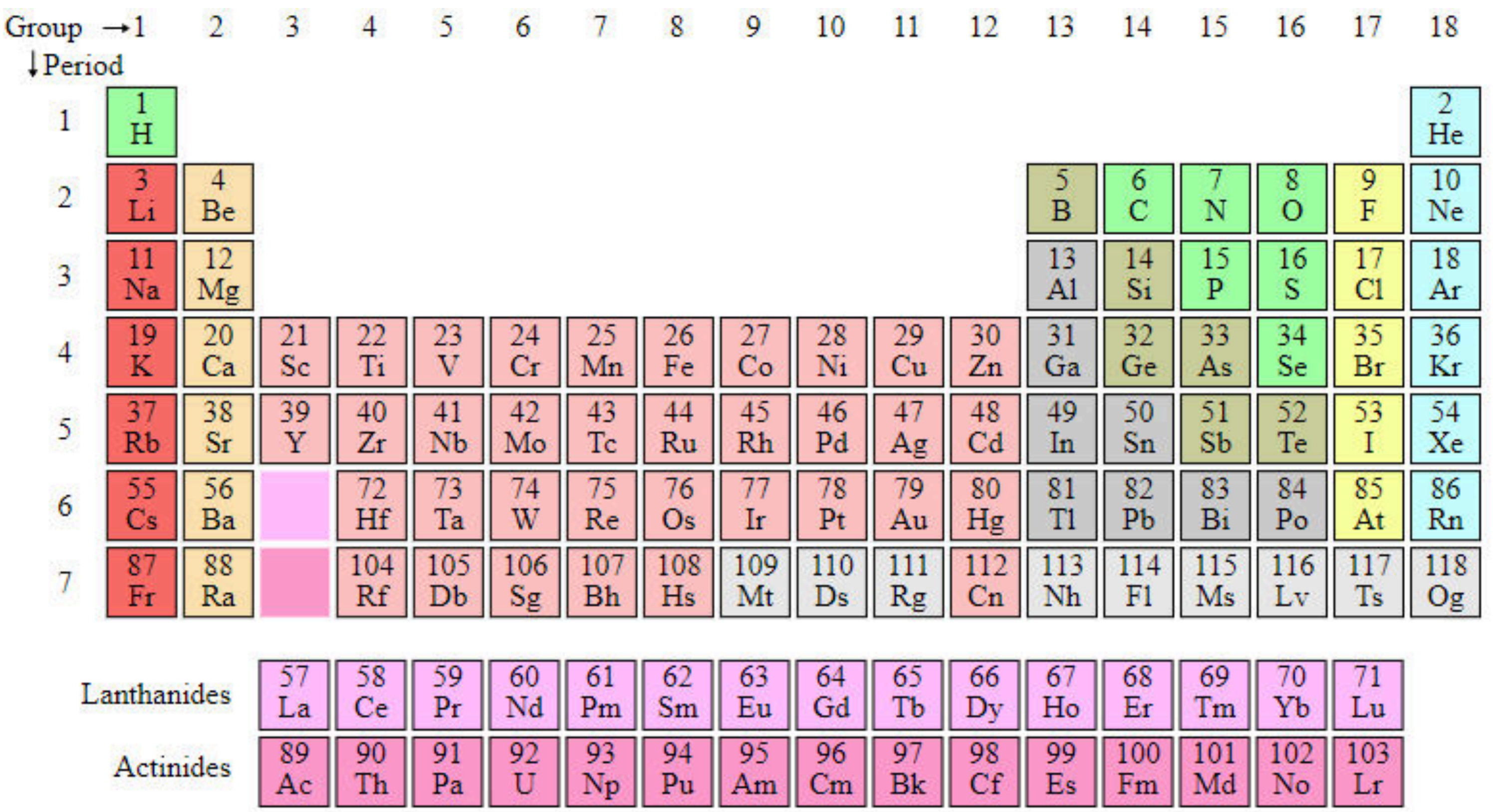


Which atom is the most electronegative? ____

1.14.2

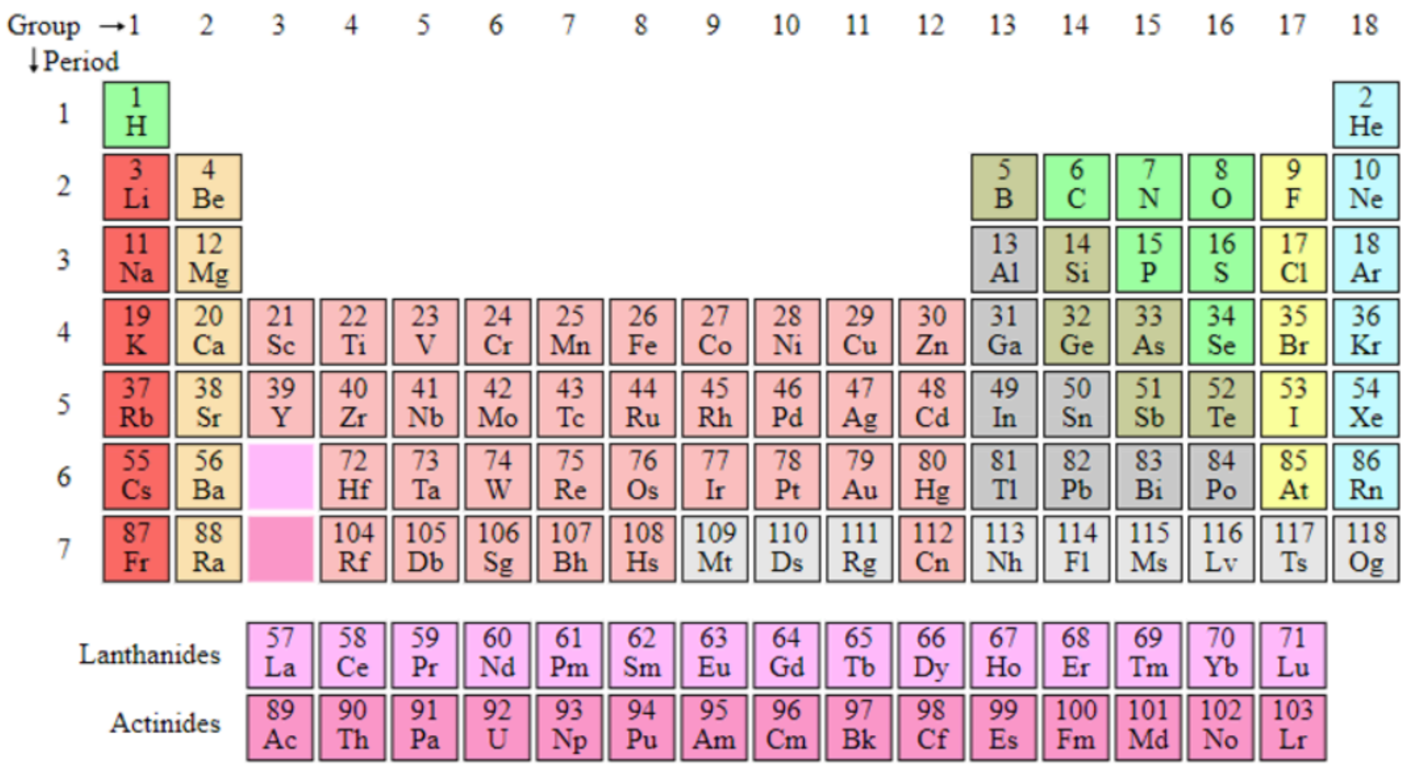
Example: Increasing Electronegativity

Rank the following atoms in increasing electronegativity: O, Al, Sr, N, Si, Cs



Practice: Ranking Electronegativity

Choose the option which correctly ranks the following elements in terms of electronegativity.

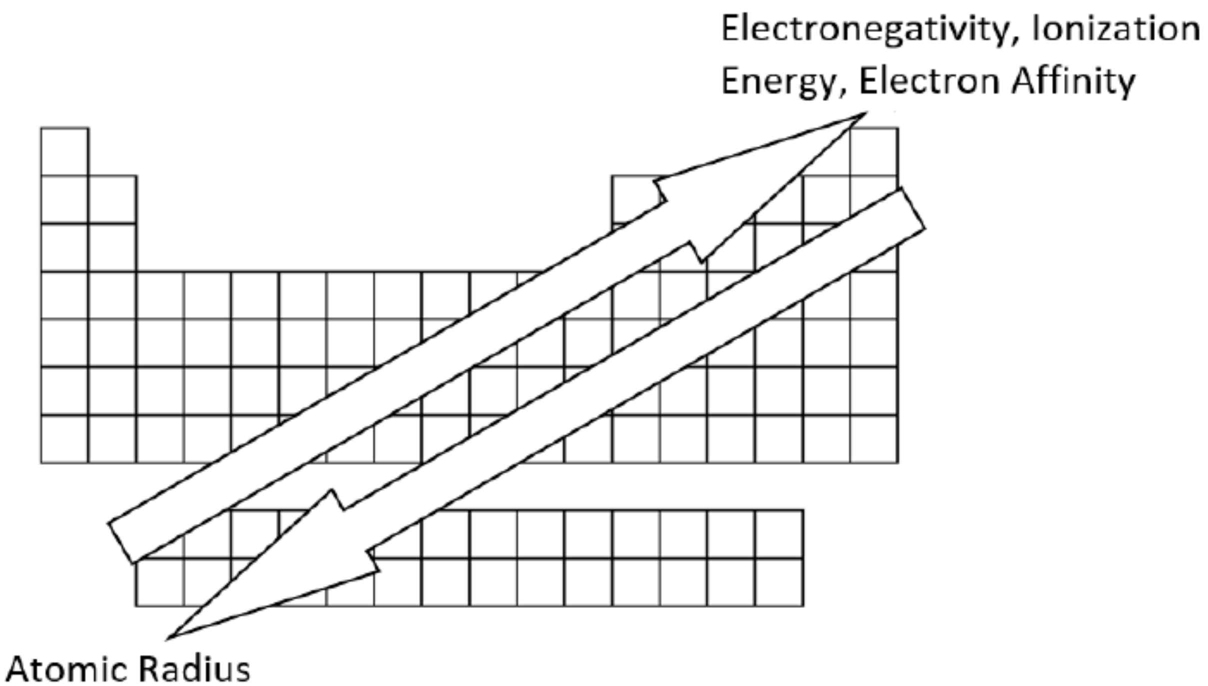


- Ti < Bi < Hg < Cs < Fr
- Na < Al < O < S < Cl
- Mg < P < C < O < F
- Fe < Co < Mo < F < Cl

1.15 Summary of Periodic Table Trends

1.15.1

Summary of Periodic Trends



Review: which direction does Zeff increase in? ____________________

1.15.2

Example: Periodic Trends

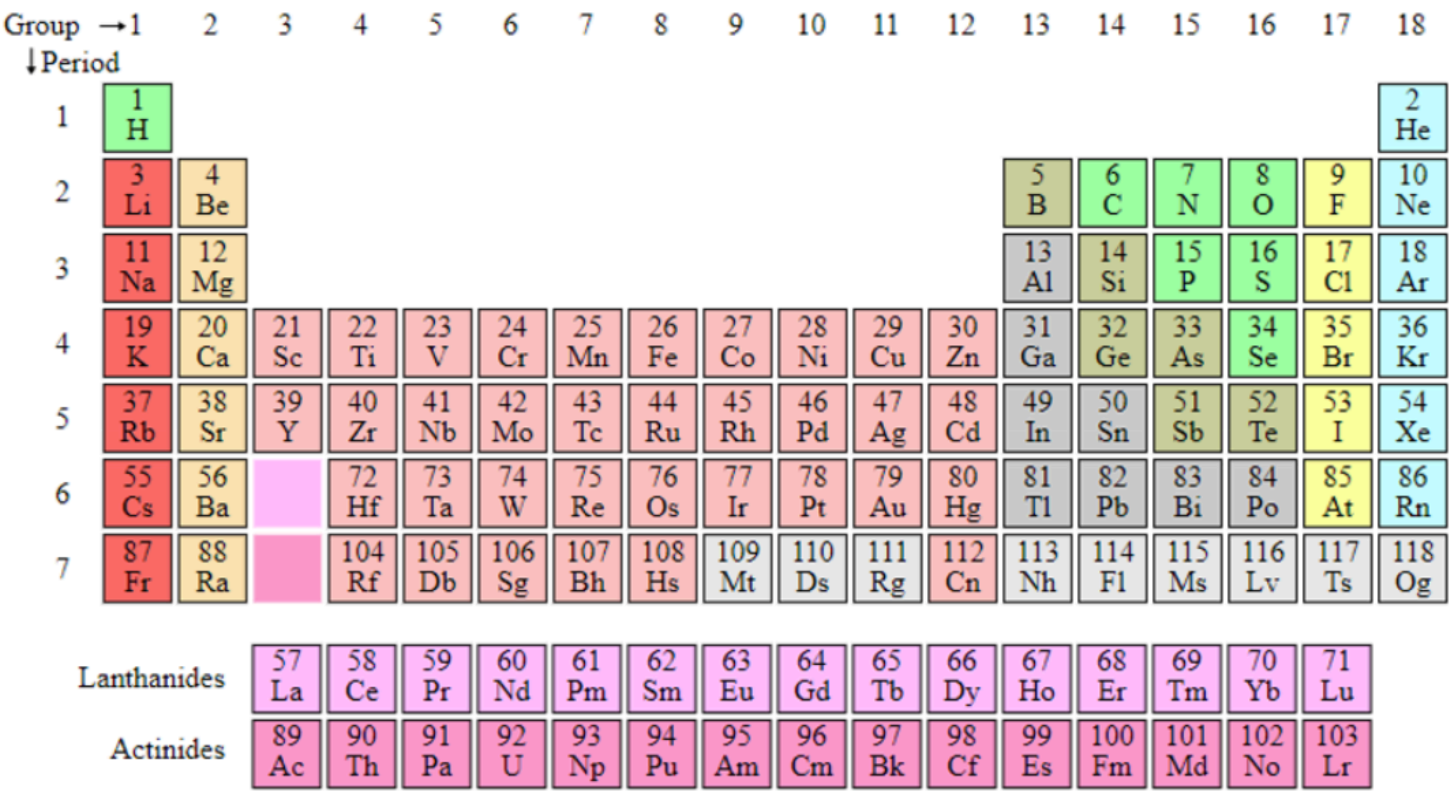
Label the following statements as either TRUE or FALSE

1. Ionization energy decreases when the atomic size decreases
2. As atomic size increases it gets easier to add an additional electron
3. Mg^{2+} is the same size as Ne since they are isoelectronic

1.15.3

Practice: Periodic Trends

Complete the statements below by filling in the blanks and then choose which option below would help explain your answer.



Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

Part 1

N^{3-} has a (larger/smaller) ionic radius in comparison to Si^{4+} because ...

- ... of the difference in Z*
- ... of the difference in n
- ... it has a higher proton to electron ratio
- ... it has a higher electron to proton ratio

Practice: Periodic Trends

Complete the statements below by filling in the blanks and then choose which option below would help explain your answer.

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

Part 2

Al has a (more negative/ less negative) EA than Si because ...

- ... of the difference in Zeff
- ... of the difference in n
- ... it has a higher proton to electron ratio
- ... it has a higher electron to proton ratio

1.15.4

Practice: Periodic Trends

Which of the following statements is correct about chlorine and phosphorus atoms?

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
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Lanthanides	57 La	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	89 Ac	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

A) The electron affinity of a chlorine atom is smaller than that of a phosphorus atom due to its smaller radius.

B) A chloride anion has higher electron affinity than a neutral chlorine atom.

C) A chlorine atom can easily gain and/or lose an electron since it has both a large electron affinity and a large ionization energy.

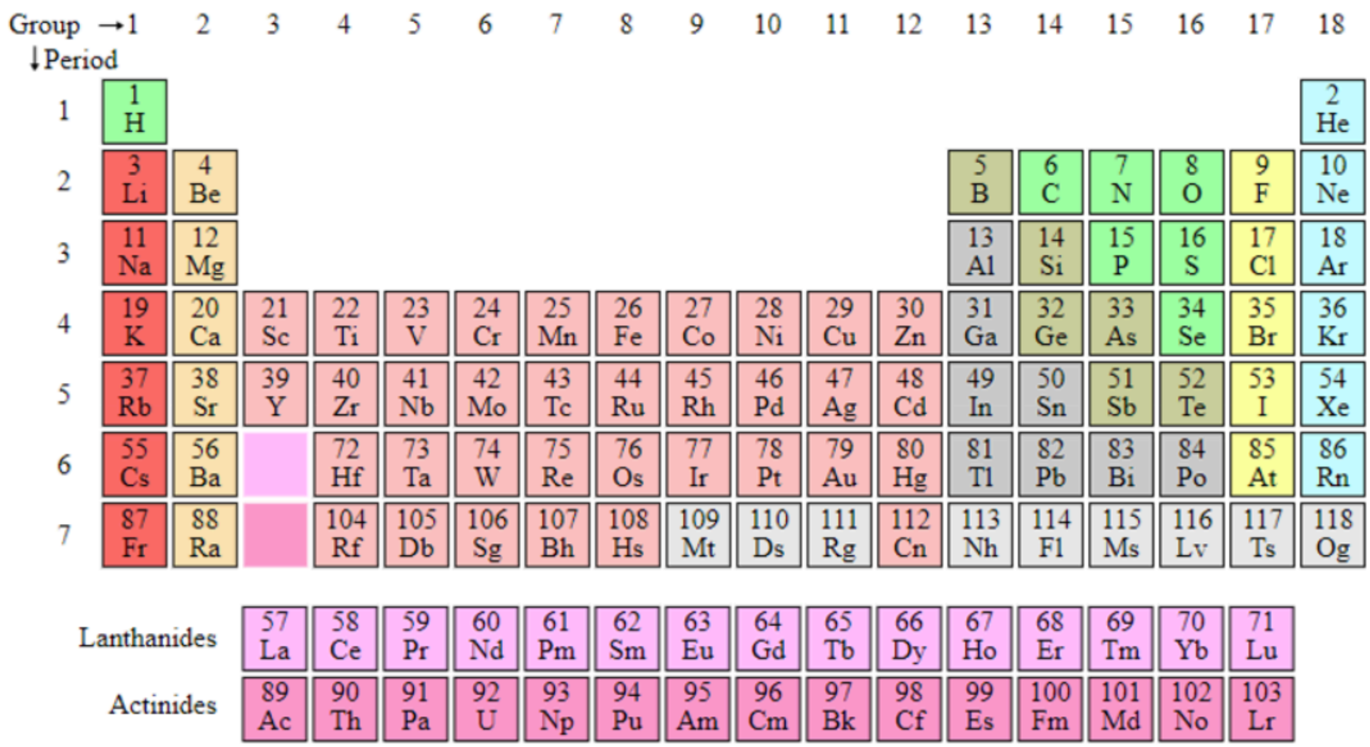
D) The effective nuclear charge experienced by the valence electrons in a chlorine atom is smaller than in a phosphorus atom.

E) All of the above statements are false.

1.15.5

Practice: Attracting Electrons

For an S-O bond, which atom will attract the electrons within the bond more strongly? Why?



- Sulfur, because it has a higher electronegativity
- Oxygen, because it has a higher electronegativity
- Sulfur, because it has a higher Zeff
- Oxygen, because it has more valence electrons
- Sulfur, because it has more protons
- Oxygen, because it is smaller

1.16 The Periodic Table & Valence Electrons

1.16.1

The Periodic Table of Elements

The **periodic table** is something we are going to see a lot of in chemistry! It organizes the elements by their **atomic number (Z)** and is organized into **groups (columns)** and **periods (rows)**.

Elements in the **same group have very similar reactivity** which we will talk about more when we learn about things like valence electrons, bonding, and Lewis structures.

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
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You should be familiar with each of the following labels:

- Groups
- Periods
- Alkali Metals
- Alkaline Earth Metals
- Transition Metals
- Nitrogen group (aka Pnictogens)
- Oxygen group (aka Chalcogens)
- Halogens
- Noble (Inert) Gases
- Metals
- Non-metals
- Metalloids
- Lanthanides and Actinides (aka Rare Earth Metals)

1.16.2

Practice: Understanding Noble Gases

Which of the following statements is true?

- Noble gases are highly reactive since they want to obtain a full octet.
- Noble gases are highly reactive since they want to gain more electrons.
- Noble gases are unreactive since they already have a full octet.
- Noble gases are very stable because they want to gain more electrons