EXTENSION ACTIVITY
GENERAL CHEMISTRY

IONIZATION ENERGY - KEY

Activity Directions

This activity will serve as practice for the topics covered in the Ionization Energy game, as well as help you build on many of the concepts you learned in the Radii Trends game. This activity is best used in conjunction with not only the tutorial levels, but also supplementary learning resources such as course lectures, textbook reading, etc. Questions labeled “Lock It In” are simply opportunities for you to solidify what you have accomplished in each task and help ensure you meet each objective.

1. Log into Collisions and navigate to the Ionization Energy Game.
2. Play the Tutorial levels, if you haven’t done so already.
3. Exit the levels and enter the Ionization Energy sandbox.
4. Follow all instructions as written below. Be sure to reference your course’s textbook, lecture notes, etc. as needed.

OBJECTIVE 1

Demonstrate an understanding of the relationship between effective nuclear charge and ionization energy.

A highly important concept in chemistry is the idea of ionization energy. Ionization energy describes the amount of energy required to remove an electron from a neutral atom in the gaseous state as described in the equation below. The energy required to remove the first electron is called the “first ionization energy”, while the removal of successive electrons earns the labels “second ionization energy”, “third ionization energy”, and so on. electrons earns the labels “second ionization energy”, “third ionization energy”, and so on.

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

Note that in the Ionization Energy game, ionization energy is described in “units of energy”. In reality, however, ionization energies are typically given as kilojoules per mole (kJ/mol) or electronvolts (eV).

TASK 1: Pull out the atom of each element listed in the table below from the atom bank. Record the number of units of energy required to remove the first valence electron (the first ionization energy) from each. The third column of the table will be left blank until *Task 2*.

Element	First Ionization Energy (units of energy)	Effective Nuclear Charge on Valence Electrons
Sodium (Na)	5	2.507
Aluminum (Al)	6	4.066
Selenium (Se)	10	8.287
Bromine (Br)	12	9.028

OBJECTIVE 1

Demonstrate an understanding of the relationship between effective nuclear charge and ionization energy.

TASK 2: The values in the periodic table below represent the effective nuclear charge on different electrons in the ground state atom of each element. Find the effective nuclear charge on **the last valence electron level** for each of the elements that you used in *Task 1* and complete the table.

	H	He
Z	1	2
1s	1.000	1.688

	Li	Be	B	C	N	O	F	Ne
Z	3	4	5	6	7	8	9	10
1s	2.691	3.685	4.680	5.673	6.665	7.658	8.650	9.642
2s	1.279	1.912	2.576	3.217	3.847	4.492	5.128	5.758
2p			2.421	3.136	3.834	4.453	5.100	5.758

	Na	Mg	Al	Si	P	S	Cl	Ar
Z	11	12	13	14	15	16	17	18
1s	10.626	11.609	12.591	13.575	14.558	15.541	16.524	17.508
2s	6.571	7.392	8.214	9.020	9.825	10.629	11.430	12.230
2p	6.802	7.826	8.963	9.945	10.961	11.977	12.993	14.008
3s	2.507	3.308	4.117	4.903	5.642	6.367	7.068	7.757
3p			4.066	4.285	4.886	5.482	6.116	6.764

	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Z	19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
1s	18.490	19.473	20.457	21.441	22.426	23.414	24.396	25.381	26.367	27.353	28.339	29.325	30.309	31.294	32.278	33.262	34.247	35.232
2s	13.006	13.776	14.574	15.377	16.181	16.984	17.794	18.599	19.405	20.213	21.020	21.828	22.599	23.365	24.127	24.888	25.643	26.398
2p	15.027	16.041	17.055	18.065	19.073	20.075	21.084	22.089	23.092	24.095	25.097	26.098	27.091	28.082	29.074	30.065	31.056	32.047
3s	8.680	9.602	10.340	11.033	11.709	12.368	13.018	13.676	14.322	14.961	15.594	16.219	16.996	17.790	18.596	19.403	20.219	21.033
3p	7.726	8.658	9.406	10.104	10.785	11.466	12.109	12.778	13.435	14.085	14.731	15.369	16.204	17.014	17.850	18.705	19.571	20.434
4s	3.495	4.398	4.632	4.817	4.981	5.133	5.283	5.434	5.576	5.711	5.842	5.965	7.067	8.044	8.944	9.758	10.553	11.316
3d			7.120	8.141	8.983	9.757	10.528	11.180	11.855	12.530	13.201	13.878	15.093	16.251	17.378	18.477	19.559	20.626
4p													6.222	6.780	7.449	8.287	9.028	9.338

	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Z	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
1s	36.208	37.191	38.176	39.159	40.142	41.126	42.109	43.092	44.076	45.059	46.042	47.026	48.010	48.992	49.974	50.957	51.939	52.922
2s	27.157	27.902	28.622	29.374	30.125	30.877	31.628	32.380	33.155	33.883	34.634	35.386	36.124	36.859	37.595	38.331	39.067	39.803
2p	33.039	34.030	35.003	35.993	36.982	37.972	38.941	39.951	40.940	41.930	42.919	43.909	44.898	45.885	46.873	47.860	48.847	49.835
3s	21.843	22.664	23.552	24.362	25.172	25.982	26.792	27.601	28.439	29.221	30.031	30.841	31.631	32.420	33.209	33.998	34.787	35.576
3p	21.303	22.168	23.093	23.846	24.616	25.474	26.384	27.221	28.154	29.020	29.809	30.692	31.521	32.353	33.184	34.009	34.841	35.668
4s	12.388	13.444	14.264	14.902	15.283	16.096	17.198	17.656	18.582	18.986	19.865	20.869	21.761	22.658	23.544	24.408	25.297	26.173
3d	21.679	22.726	25.397	25.567	26.247	27.228	28.353	29.359	30.405	31.451	32.540	33.607	34.678	35.742	36.800	37.839	38.901	39.947
4p	10.881	11.932	12.746	13.460	14.084	14.977	15.811	16.435	17.140	17.723	18.562	19.411	20.369	21.265	22.181	23.122	24.030	24.957
5s	4.985	6.071	6.256	6.446	5.921	6.106	7.227	6.485	6.640	(empty)	6.756	8.192	9.512	10.629	11.617	12.538	13.404	14.218
4d			15.958	13.072	11.238	11.392	12.882	12.813	13.442	13.618	14.763	15.877	16.942	17.970	18.974	19.960	20.934	21.893
5p													8.470	9.102	9.995	10.809	11.612	12.425

Effective Nuclear Charge values from Clementi et al. 1963 and 1967

LOCK IT IN: Based on what you see with your atoms, what broad relationship exists between the effective nuclear charge and the first ionization energy?

Elements whose valence electrons experience larger effective nuclear charges tend to have higher ionization energies.

OBJECTIVE 2

Demonstrate an understanding of the successive ionization energies.

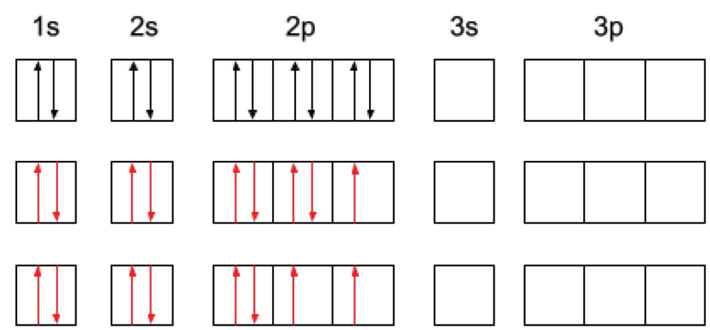
TASK 3: For this task you will explore successive ionization energies. The sandbox allows you to remove more than just one electron from an atom. As such, you are able to see how the amount of energy required to remove successive electrons changes.

1. In the designated spaces below, enter the first, second, and third ionization energies for sodium and magnesium determined using the sandbox.
2. You must also determine the difference/size of the jump between the first and second and second and third ionization energies.
3. To aid in your understanding of what each ionization energy represents, complete the electron configuration and orbital diagrams of the atom after each electron removal.
4. The first few sections have been completed for the sodium atom to help you.

Sodium

	Ionization Energy (units of energy)	Difference From Previous IE (units of energy)	Electron Configuration
First IE	5		$1s^2 2s^2 2p^6 3s^0$
Second IE	46	39	$1s^2 2s^2 2p^5 3s^0$
Third IE	70	34	$1s^2 2s^2 2p^4 3s^0$

Orbital Diagram



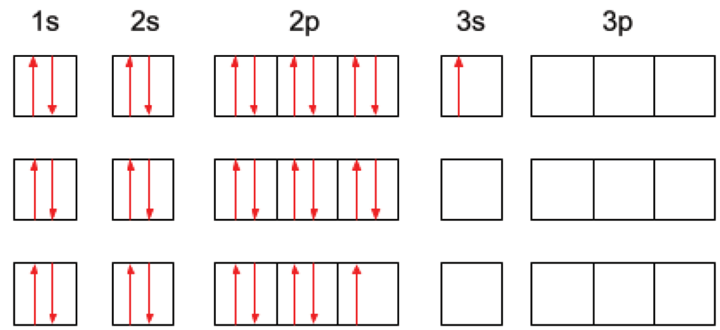
OBJECTIVE 2

Demonstrate an understanding of the successive ionization energies.

Magnesium

	Ionization Energy (units of energy)	Difference From Previous IE (units of energy)	Electron Configuration
First IE	8		$1s^22s^22p^63s^1$
Second IE	15	7	$1s^22s^22p^63s^0$
Third IE	78	63	$1s^22s^22p^53s^0$

Orbital Diagram



LOCK IT IN:
Identify the trend in successive ionization energies. Explain why this trend occurs.

LOCK IT IN:
Explain why the largest jump in ionization energy does not occur at the same point in sodium and magnesium. Hint: Look at the orbital diagram to see where the electron being removed is located.

Ionization energies increase with each successive ionization. This occurs because each time an electron is removed, it is removed from an increasingly net positive ion.

The largest jump in ionization energy does not occur at the same point in magnesium and sodium because they have different numbers of valence electrons. The removal of the second electron in sodium requires breaking the octet of the second energy level (which requires a lot of energy) because sodium only has one valence electron. However, magnesium has two valence electrons, so breaking the octet of the second energy level does not happen until the removal of the third electron.

OBJECTIVE 3

Demonstrate an understanding of exceptions to normal patterns of ionization energy.

TASK 4: We can use the sandbox and our understanding of electron configurations to study some special phenomena involving ionization energies. The first of these involves the elements beryllium and boron.

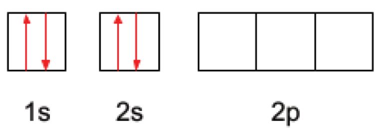
To begin, look at the effective nuclear charges of beryllium (Be) and boron (B) using the table from *Task 2*. As you remember, effective nuclear charge increases from left to right across a period. Boron is located just to the right of beryllium in the same period. In the space below, write a claim as to which of the two should have the higher first ionization energy and provide your reasoning.

Boron should theoretically have the higher ionization energy due to its position to the right of beryllium in the same period. This tells us that the valence electrons experience a higher effective nuclear charge in boron than in beryllium.

In the sandbox, remove a beryllium atom and a boron atom from the bank. Complete the orbital diagram for each atom before removing one electron and recording the ionization energy in the spaces below. After you remove the electron, complete the orbital diagram for each ion.

Beryllium

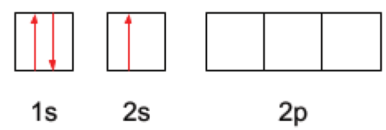
Initial Orbital Diagram



First Ionization Energy
(units of energy)

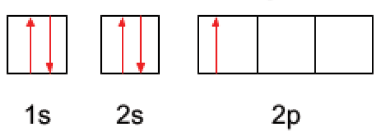
9

Final Orbital Diagram



Boron

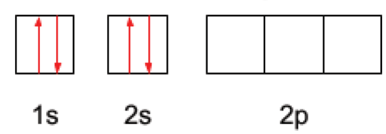
Initial Orbital Diagram



First Ionization Energy
(units of energy)

8

Final Orbital Diagram



OBJECTIVE 3

Demonstrate an understanding of exceptions to normal patterns of ionization energy.

Did the ionization energies modeled in the sandbox support or refute your initial claim about the first ionization energies of beryllium and boron? Use the data you collected to support your decision.

The ionization energies from the sandbox refute the initial claim that was made. Boron has a first ionization energy of 8 units of energy compared to 9 in beryllium. This goes against the claim that boron would have the larger first ionization energy.

LOCK IT IN:
When thinking about atoms and ions, it is very important to remember that there is a very complex balance of attractive and repulsive forces between electrons and protons and electrons with other electrons. These interactions are complemented by a complex suite of other factors and aspects of quantum mechanics. One result of these different effects is that fully filled subshells are more stable than partially filled ones. Use this information to explain why you observed what you did when determining the first ionization energies of beryllium and boron.

The reason beryllium had an unexpectedly higher first ionization energy than boron is because removing its first valence electron would require breaking up a filled 2s orbital. However, removing the first electron from boron required removing an electron from the otherwise empty 2p subshell. In this case, the filled 2s subshell is more stable than the partially filled 2p subshell.

OBJECTIVE 3 Demonstrate an understanding of exceptions to normal patterns of ionization energy.

TASK 5: We will now explore another ionization energy phenomenon using the elements nitrogen and oxygen.

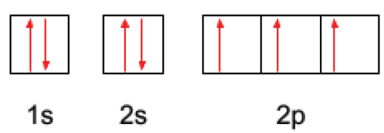
To begin, look at the effective nuclear charges of nitrogen (N) and oxygen (O) using the table from *Task 2*. As you remember, effective nuclear charge increases from left to right across a period. Oxygen is located just to the right of nitrogen in the same period. In the space below, write a claim as to which of the two should have the higher first ionization energy and provide your reasoning.

Oxygen should theoretically have the higher ionization energy due to its position to the right of nitrogen in the same period. This tells us that the valence electrons experience a higher effective nuclear charge in oxygen than in nitrogen.

In the sandbox, remove a nitrogen atom and an oxygen atom from the bank. Complete the orbital diagram for each atom before removing one electron and recording the ionization energy in the spaces below. After you remove the electron, complete the orbital diagram for each ion.

Nitrogen

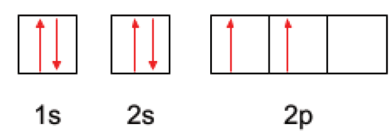
Initial Orbital Diagram



First Ionization Energy (units of energy)

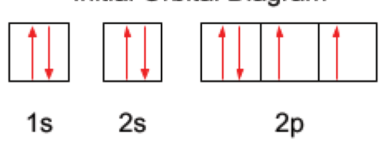
15

Final Orbital Diagram



Oxygen

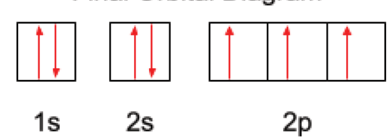
Initial Orbital Diagram



First Ionization Energy (units of energy)

14

Final Orbital Diagram



OBJECTIVE 3

Demonstrate an understanding of exceptions to normal patterns of ionization energy.

Did the ionization energies modeled in the sandbox support or refute your initial claim about the first ionization energies of nitrogen and oxygen? Use the data you collected to support your decision.

The ionization energies from the sandbox refute the initial claim that was made. Oxygen has a first ionization energy of 14 units of energy compared to 15 in nitrogen. This goes against the claim that oxygen would have the larger first ionization energy.

LOCK IT IN:
Another result of the complexity of the atom is that exactly half-filled subshells are also more stable than partially filled ones. Use this information to explain why you observed what you did when determining the first ionization energies of nitrogen and oxygen.

The reason nitrogen had an unexpectedly higher first ionization energy than oxygen is because removing its first valence electron would require breaking up a half-filled 2p subshell. However, removing the first electron from oxygen required removing the fourth electron from 2p subshell and thus creating the more stable half-filled 2p configuration. In this case, removing oxygen's first valence electron reduces repulsion and is thus energetically more favorable than breaking up the half-filled 2p subshell of nitrogen.

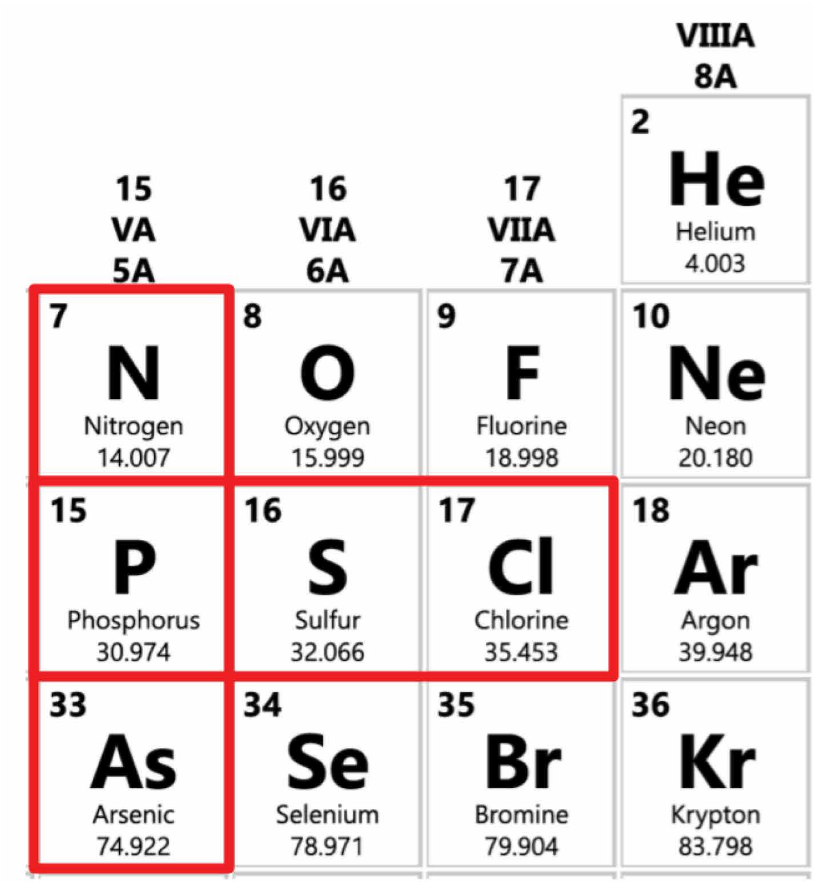
LOCK IT IN:
Describe in a few sentences the relationship between effective nuclear charge and ionization energy going across a period. Explain why there are occasionally exceptions to this relationship.

In general, as you move from left to right across a period, the attraction of the nucleus on the valence electrons (effective nuclear charge) increases and thus also increases the amount of energy required to remove an electron (ionization energy). However, due to the stability of filled and half-filled subshells, exceptions to this standard pattern arise. This is best shown as one moves from Group 2 to Group 3 elements or Group 5 to Group 6 elements.

OBJECTIVE 4

Demonstrate an understanding of the trends in ionization energy on the periodic table.

TASK 6: Use the sandbox to identify the ionization energies of the elements indicated on the section of the periodic table below.

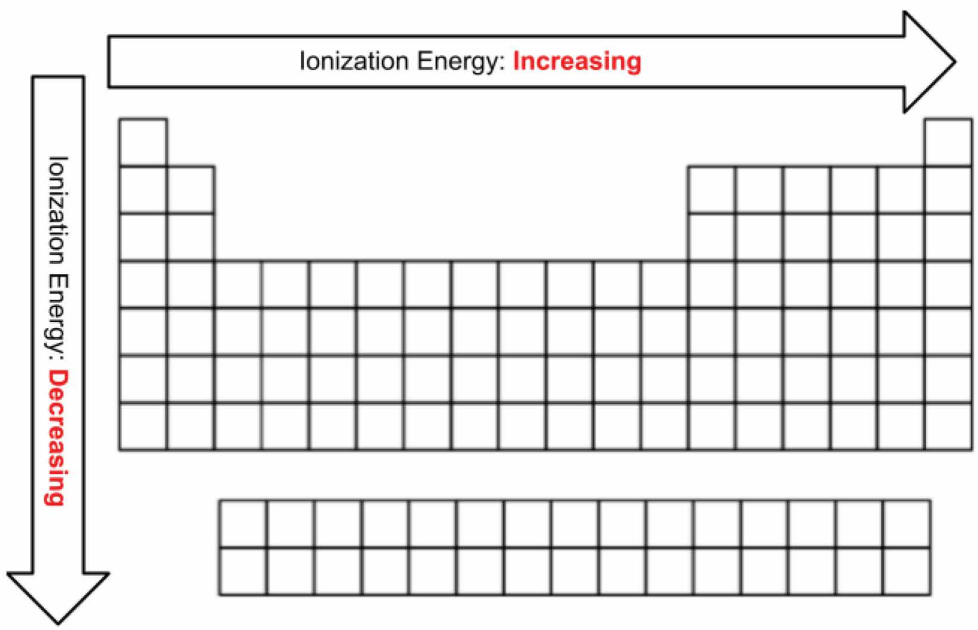


Element	First Ionization Energy (units of energy)
Nitrogen	15
Phosphorus	11
Arsenic	10
Sulfur	10
Chlorine	13

OBJECTIVE 4

Demonstrate an understanding of the trends in ionization energy on the periodic table.

LOCK IT IN:
Label the periodic table below with the terms "increasing" or "decreasing" based on the general trends in ionization energy. Remember that there will be exceptions!



LOCK IT IN:
Justify the trend you identified going from left to right across a period using the concept of effective nuclear charge.

As you move from left to right across a period, the effective nuclear charge on the valence electrons increases and thus it requires more energy to remove an electron.

LOCK IT IN:
Justify the trend you identified going down a group using the concept of increasing principal quantum number (n).

As you go down a group, the principal quantum number increases and thus so does the size of the atom. This increase in size means that valence electrons will spend much of their time far from the nucleus and thus be easier to remove.

OBJECTIVE 5

Demonstrate an understanding of the trends in electron affinity on the periodic table.

Another important concept in chemistry is the idea of electron affinity. Electron affinity describes the energy change associated with adding an electron to a neutral atom in the gaseous state as described in the equation below.

$$X(g) + e^- \rightarrow X^-(g)$$

Note that in the Ionization Energy game, electron affinity is described in “units of energy”. In reality, however, electron affinities are typically given as kilojoules per mole (kJ/mol) or electronvolts (eV). Unlike ionization energy, electron affinities can either be negative, positive, or zero. A negative electron affinity indicates that energy is released by the addition of the electron, while a positive one indicates that energy is required to add the electron. Accordingly, a very negative electron affinity can be described as a very high electron affinity while the opposite is true for a very positive electron affinity.

TASK 7: Pull out the atom of each element listed in the table below from the atom bank. Record the number of units of energy released (include a negative sign) or used (include a positive sign) to add an electron to each element.

Element	Electron Affinity (units of energy)
Sodium (Na)	- 1
Magnesium (Mg)	+ 1
Aluminum (Al)	- 1
Sulfur (S)	- 2
Chlorine (Cl)	- 4

OBJECTIVE 5

Demonstrate an understanding of the trends in electron affinity on the periodic table.

TASK 8: Trends in electron affinity are less identifiable going down a group of the periodic table. However, there are other patterns that emerge, particularly across a period. Use the information you gathered in Task 6 and the image of Period 3 below to help you answer the questions beneath it.

Metals				Nonmetals			
sodium 11 Na 22.990	magnesium 12 Mg 24.305	aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948

LOCK IT IN:
What is the general trend in electron affinity that exists going from left to right across a period?

Electron affinity generally increases from left to right across a period.

LOCK IT IN:
Argon (Ar) has by the far the lowest electron affinity of the elements in period 7. Explain why this makes sense using its electron configuration of $1s^22s^22p^63s^23p^6$.

Argon already has a complete octet and thus it is least energetically favorable to add an electron to an atom of argon.

LOCK IT IN:
Assuming the data you collected in *Task 7* is representative (which it is), compare the electron affinities of metals with nonmetals (excluding argon).

Metals typically have lower electron affinities than nonmetals.

LOCK IT IN:
Compare the electron affinity you observed for sodium with that of magnesium. Explain why that makes sense considering that the electron configuration of sodium is $1s^22s^22p^63s^1$ and the electron configuration of magnesium is $1s^22s^22p^63s^2$.

Adding an electron to sodium simply completes a 3s orbital, while adding an electron to a magnesium atom requires adding it into a new 3p orbital.

OBJECTIVE 6

Demonstrate an understanding of how the octet rule helps determine the charge of a main group ion.

One of the most crucial concepts in understanding the main group elements and their behaviors is the octet rule. This phenomenon describes the tendency of atoms to prefer having eight electrons in the valence shell. More specifically, atoms prefer to have full valence s and p subshells and, as you will see in other games, will react with other elements in ways that allow them to achieve such a state as best as possible. This same situation is often described as an attempt by atoms to resemble the nearest noble gas (Group 8) on the periodic table since noble gases in their neutral ground state all have full valence shells.

TASK 9: In this task, you will once again take a look at the elements of Period 3. However, this time you will be determining how to help these elements satisfy the octet rule in the most energetically favorable manner.

1. Use the sandbox to determine the amount of energy required to reach a complete octet by removing electrons from each atom and then by adding electrons. In the case that there is not enough energy available in the sandbox to complete a task, simply enter " >100" to indicate that more than 100 units of energy would be necessary to complete the removal or addition of an electron.
2. Decide if it is more energetically favorable to add or remove electrons and indicate how many electrons should be added or removed.
3. Determine the charge of the ion satisfying the octet rule. Remember that electrons are negatively charged!

Element	Energy Required to Add Electrons to Complete Octet (units of energy)	Energy Required to Remove Electrons to Complete Octet (units of energy)	Most Energetically Favorable Scenario? ("gain electrons" or "lose electrons")	Change in # of Electrons for Most Energetically Favorable Scenario	Charge of Ion Satisfying Octet Rule
Sodium (Na)	>100	+ 5	lose electrons	lose 1	1+
Magnesium (Mg)	>100	+ 23	lose electrons	lose 2	2+
Aluminum (Al)	>100	+ 53	lose electrons	lose 3	3+
Phosphorus (P)	+ 13	>100	gain electrons	gain 3	3-
Sulfur (S)	+ 3	>100	gain electrons	gain 2	2-
Chlorine (Cl)	- 4	>100	gain electrons	gain 1	1-

OBJECTIVE 6

Demonstrate an understanding of how the octet rule helps determine the charge of a main group ion.

Metals				Nonmetals			
sodium	magnesium	aluminium	silicon	phosphorus	sulfur	chlorine	argon
11	12	13	14	15	16	17	18
Na	Mg	Al	Si	P	S	Cl	Ar
22.990	24.305	26.982	28.086	30.974	32.065	35.453	39.948

LOCK IT IN:
According to trends in the Period 3 elements, do **metals** generally prefer to gain or lose electrons to achieve a complete octet?

Metals generally prefer to lose electrons to complete an octet.

LOCK IT IN:
Using what you have seen so far in this sandbox activity, is it more energetically favorable to form cations from elements with low ionization energies to complete the octet rule or from those with high ionization energies?

It is generally more energetically favorable to form cations from elements with low ionization energies.

LOCK IT IN:
According to trends in the Period 3 elements, do **nonmetals** generally prefer to gain or lose electrons to achieve a complete octet?

Nonmetals generally prefer to gain electrons to complete an octet.

LOCK IT IN:
Using what you have seen so far in this sandbox activity, is it more energetically favorable to form anions from elements with low electron affinities to complete the octet rule or from those with high electron affinities?

It is generally more energetically favorable to form anions from elements with high electron affinities.

OBJECTIVE 7

Demonstrate an understanding of how the ionic radius is determined by the difference between the nuclear charge and the number of electrons.

TASK 10: Complete the table with your knowledge of electron configurations and atomic radii gained during the Radii Trends game. Then rank the atoms by radius in the designated space beneath the table. Use a periodic table as necessary.

Element	# of Protons	# of Electrons	Electron Configuration	Orbital Diagram
Sulfur (S)	16	16	$1s^2 2s^2 2p^6 2s^2 3s^2 3p^4$	1s ↑↓ · 2s ↑↓ · 2p ↑↓ ↑↓ ↑↓ · 3s ↑↓ · 3p ↑↓ ↑ ↑ · 4s
Chlorine (Cl)	17	17	$1s^2 2s^2 2p^6 2s^2 3s^2 3p^5$	1s ↑↓ · 2s ↑↓ · 2p ↑↓ ↑↓ ↑↓ · 3s ↑↓ · 3p ↑↓ ↑↓ ↑ · 4s
Potassium (K)	19	19	$1s^2 2s^2 2p^6 2s^2 3s^2 3p^6 4s^1$	1s ↑↓ · 2s ↑↓ · 2p ↑↓ ↑↓ ↑↓ · 3s ↑↓ · 3p ↑↓ ↑↓ ↑↓ · 4s ↑
Calcium (Ca)	20	20	$1s^2 2s^2 2p^6 2s^2 3s^2 3p^6 4s^2$	1s ↑↓ · 2s ↑↓ · 2p ↑↓ ↑↓ ↑↓ · 3s ↑↓ · 3p ↑↓ ↑↓ ↑↓ · 4s ↑↓

Potassium	**Largest Radius**
Calcium	
Sulfur	
Chlorine	**Smallest Radius**

OBJECTIVE 7

Demonstrate an understanding of how the ionic radius is determined by the difference between the nuclear charge and the number of electrons.

TASK 11: Create the ions listed in the table below in the sandbox and make sure to “Check” them so that they appear in the bottom panel. Once they all appear there, complete the rest of the table and rank them by ionic radius.

Ion	# of Protons	# of Electrons	Electron Configuration	Orbital Diagram
S^{2-}	16	18	$1s^22s^22p^62s^23s^23p^6$	1s [↑↓] 2s [↑↓] 2p [↑↓][↑↓][↑↓] 3s [↑↓] 3p [↑↓][↑↓][↑↓] 4s [ ]
Cl^-	17	18	$1s^22s^22p^62s^23s^23p^6$	1s [↑↓] 2s [↑↓] 2p [↑↓][↑↓][↑↓] 3s [↑↓] 3p [↑↓][↑↓][↑↓] 4s [ ]
K^+	19	18	$1s^22s^22p^62s^23s^23p^6$	1s [↑↓] 2s [↑↓] 2p [↑↓][↑↓][↑↓] 3s [↑↓] 3p [↑↓][↑↓][↑↓] 4s [ ]
Ca^{2+}	20	18	$1s^22s^22p^62s^23s^23p^6$	1s [↑↓] 2s [↑↓] 2p [↑↓][↑↓][↑↓] 3s [↑↓] 3p [↑↓][↑↓][↑↓] 4s [ ]

S^{2-}	**Largest Radius**
Cl^-	
K^+	
Ca^{2+}	**Smallest Radius**

OBJECTIVE 7

Demonstrate an understanding of how the ionic radius is determined by the difference between the nuclear charge and the number of electrons.

LOCK IT IN:
Are the trends in atomic radius and ionic radius the same? Explain your answer.

The trends in atomic and ionic radii are not the same. While sulfur and chlorine atoms are smaller than potassium and calcium atoms, sulfur and chlorine ions are larger than those of potassium or calcium.

LOCK IT IN:
You should notice that the ions in Task 10 are isoelectronic—ions with the same electron configurations. Explain why these ions do not all have the same radius. As part of your answer, justify why the smallest and largest ions are that way.

Despite being isoelectronic, the ions in Task 10 have different numbers of protons and thus their electrons are being pulled towards the nucleus by different amounts of charge. The largest ion, S2-, has two more electrons than it does protons and thus its electrons can move away farther from the nucleus. The smallest ion, Ca2+, has two more protons than electrons, which means it will draw its electrons in much closer.

LOCK IT IN:
Compare the radius of cations and anions with their parent atoms.

Anions become larger than their parent atoms, while cations become smaller than their parent atoms.

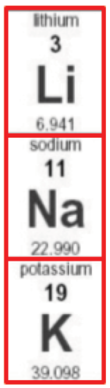
LOCK IT IN:
With which noble gas do these elements all share an electron configuration?

Argon

OBJECTIVE 8

Demonstrate an understanding of the trends in ionic radius on the periodic table.

TASK 12: Depending on what ions remain in the bottom panel of your sandbox, ensure that you either have or create lithium (Li^+), sodium (Na^+), and potassium (K^+) ions. Rank them by radius in the space below.

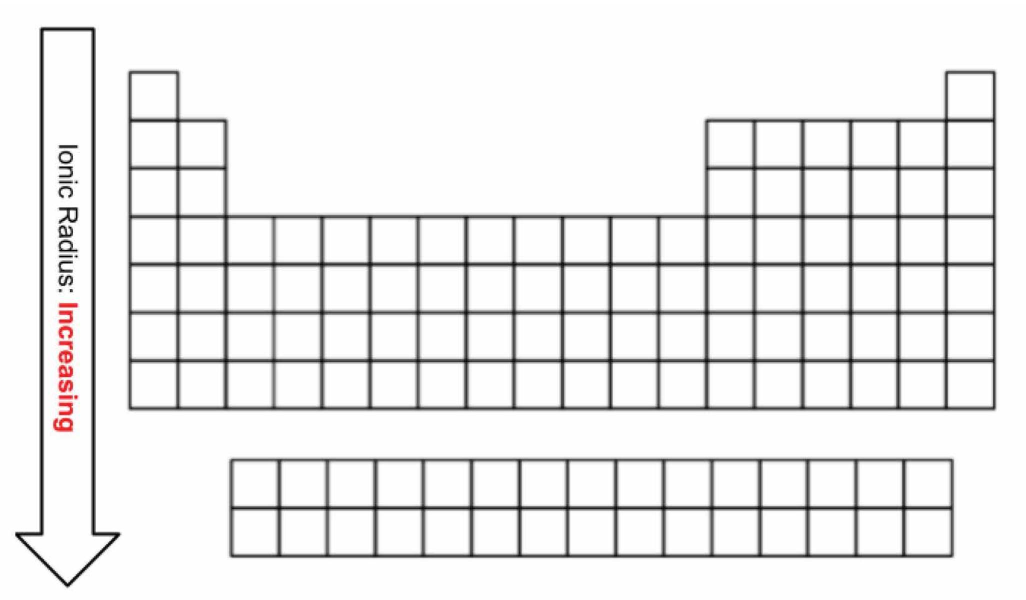


K^+	**Largest Radius**
Na^+	
Li^+	**Smallest Radius**

OBJECTIVE 8

Demonstrate an understanding of the trends in ionic radius on the periodic table.

LOCK IT IN:
The trend in ionic radius going right across a period for the main group elements (transition elements are more complex) is that the radius decreases for the positive ions and then increases at the first negative ion and then decreases from there. Label the periodic table below with the term "increasing" or "decreasing" based on the trend in ionic radius going down a group.



CLOSURE

CLOSURE: Rubidium (Rb) and Iodine (I) are two elements in Period 5 that are not available to you in the sandbox for the Ionization Energy game. However, you should be able to use what you have learned so far to demonstrate your overall understanding of the concepts presented in the game. Compare the ions for rubidium and iodine on their ionization energy, electron affinity, and ionic radius using only a periodic table. Enter a greater than (>) or less than (<) symbol into the table. Then provide a brief justification as to why you chose the symbol you did using what you have learned.

In the table, identify how many electrons the atom would lose or gain to satisfy the octet rule and identify the charge of the resulting ion. You must also justify your answers there as well.

	53 I	> or <	37 Rb	Justification
Ionization Energy		>		Iodine is further to the right than rubidium in period 5.
Electron Affinity		>		Iodine is further to the right than rubidium in period 5.
Ionic Radius		>		Iodine has seven valence electrons and will gain an electron, while rubidium will lose its one.

Element	# of electrons Lost or Gained to Satisfy Octet Rule	Charge of Resulting Ion	Justification
Rubidium	lose one	1+	Rubidium is a metal in Group 1, so it is most energetically favorable to lose its one valence electron to complete its octet.
Iodine	gain one	1-	Iodine is a nonmetal in Group 7, so it is most energetically favorable to gain one valence electron to complete its octet.