

Section 11: Electron Configuration and Periodic Trends

The following maps the videos in this section to the Texas Essential Knowledge and Skills for Science TAC §112.35(c).

11.01 The Bohr Model of the Atom

- Chemistry (6)(A)

11.02 Orbital Energies and Electron Configurations

- Chemistry (6)(E)

11.03 Condensed Electron Configurations and Exceptions

- Chemistry (6)(E)

11.04 Categories of Electrons

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11.05 Trends in Atomic Radii

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11.06 Trends in Ionization Energy

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11.07 Electron Configuration for Ions

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11.08 Trends in Ionic Radii

- Chemistry (5)(C)

Note: Unless stated otherwise, any sample data is fictitious and used solely for the purpose of instruction.

11.01

The Bohr Model of the Atom

In 1913, Danish physicist **Niels Bohr** proposed a new model of the hydrogen atom. His model has the following characteristics:

- The hydrogen atom has **energy levels** (n) at which the electron can travel in a circular orbit around the nucleus. The smaller the value of n , the smaller the radius of the orbit and the lower the energy level.
- An electron in an atom does not radiate energy while orbiting the nucleus.
- The electron will move to another energy level only if it **absorbs** or **emits** a photon that has the *same energy* as the difference between the two energy levels. In other words, if an electron is at energy level 1 (n_1) and absorbs a photon, it can move to energy level 2 (n_2), assuming the photon added the same amount of energy as the difference between n_1 and n_2 ($n_2 - n_1$).

When studying the Bohr model, it is essential to know a few key terms:

- **Ground state** – when the electron is at the lowest energy orbital it can occupy ($n = 1$ for hydrogen)
- **Excited state** – when the electron is at a higher energy orbital than the ground state ($n = 2, 3, 4$, etc. for hydrogen)
- **Ionize** – the process by which an electron is ejected from an atom ($n_{final} = \infty$)
- **Emission** – a transition in an atom where an electron in a higher energy level returns to a lower energy level (for example, $n = 5$ to $n = 2$)
- **Absorption** – a transition in an atom where an electron in a lower energy level moves to a higher energy level (for example, $n = 2$ to $n = 5$)

When light from an electrically excited gaseous atom is passed through a slit and refracted by a prism, it creates a **line spectrum**. This consists of a series of fine lines at specific frequencies, separated by black spaces.

In a hydrogen atom, an electron undergoing emission can produce three series of spectral lines that are particularly noteworthy:

- The **ultraviolet** spectrum is produced where the electron transition terminates at $n = 1$. The wavelengths of the lines in this series are between 90 and 100 nm.
- The **visible light** spectrum is produced where the electron transition terminates at $n = 2$. The wavelengths of the lines in this series are between 400 and 750 nm.
- The **infrared** series is produced when the electron transition terminates at $n = 3$. The wavelengths of the lines in this series are in the range of 1000s of nanometers.

1. An electron in a hydrogen atom can undergo any of the transitions below, thus changing its energy. Which transition is associated with the ultraviolet series?

- A. $n = 4$ to $n = 3$
- B. $n = 1$ to $n = 2$
- C. $n = 4$ to $n = 2$
- D. $n = 3$ to $n = 1$

2. Of the transitions in the previous problem, which will emit the radiation with the shortest wavelength?

11.02

Orbital Energies and Electron Configurations

In a multi-electron atom, the **electron configuration** is a useful tool for determining the location of electrons relative to others as they orbit the nucleus. We can think of the electron configuration as the address of the electrons in an atom.

Four Components of the Address

- The **principal energy level** (n) describes the **shell** in which the electron orbits the nucleus, according to the Bohr model of the atom. The value of n must be a positive integer (1, 2, 3, . . .), and as the value increases, the radius of the shell grows larger.
- The different energy levels within a principal energy level are called the **subshells** or **sublevels**. The sublevels, which have varying energies and shapes, are denoted by the letters s , p , d , and f . The number of sublevels in any given principal energy level is equal to n .
- The atomic **orbital** is the three-dimensional volume within a sublevel where electrons have the highest probability of being found. Every orbital can contain a maximum of 2 electrons.
 - There is 1 s orbital, which has a **spherical** shape.
 - There are 3 p orbitals, each having a **dumbbell** shape.
 - There are 5 d orbitals. Four of the five have a **double dumbbell (clover)** shape. One d orbital has a **dumbbell with a torus (donut)** shape.
 - There are 7 f orbitals, whose shapes are much more complex than those of the d orbitals.
- The **spin** of an electron is a property that can be described as either clockwise or counter-clockwise. The **Pauli Exclusion Principle** states that to minimize repulsion, electrons in the same orbital must have opposing spins.

Energy Level (n)	Sublevels Present	Maximum Electrons
1		
2		
3		
4		

1. Which of the following statements regarding energy levels and orbitals is *false*?
- A. The number of orbitals in the $5d$ sublevel is 5.
 - B. The shape of the $3s$ orbital is spherical.
 - C. There are 3 sublevels in the $n = 3$ energy level.
 - D. The maximum number of electrons in an $n = 3$ energy level is 9.

Constructing an Electron Configuration

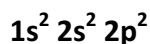
The ***aufbau principle*** states that the lowest energy orbitals in an atom are occupied first. Accordingly, when we write an electron configuration, we start filling orbitals in the lowest value energy level ($n = 1$) and build up until we have accounted for every electron in the atom.

- A neutral hydrogen atom has one electron. The lowest energy level is $n = 1$ and it contains only an s sublevel, so the electron configuration for hydrogen is



Note that the leading number refers to the n value, the letter (s) refers to the sublevel designation, and the superscript describes the number of electrons in that sublevel.

- Recall that in the $n = 2$ energy level, there are two possible sublevels, s and p . Once the s sublevel is filled, electrons begin to occupy the p sublevel. Consider the electron configuration for carbon:



- Write the electron configuration for neon, following the example given above.

- Once the s sublevel and the p sublevel are completely full, we must start on the next energy level, $n = 3$, where we can make use of s , p , and d orbitals.
- ***Hund's Rule*** states that each orbital in a sublevel is assigned one electron before any orbital is filled with a second electron.

2. Give the electron configurations for the following elements:

i. Boron

ii. Phosphorus

11.03

Condensed Electron Configurations and Exceptions

For an atom with many electrons, we can use a shortcut to express its electron configuration using the most immediate previous noble gas.

- For example, the electron configuration of calcium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$. Using the shortcut, we can express this as $[\text{Ar}] 4s^2$.
- Write the electron configuration for chlorine using the shortcut described above.

Filling Order of Electrons

The d orbitals of an energy level are filled *after* the s orbital of the next energy level. For example, the $4s$ sublevel is filled before the $3d$ sublevel.

The diagram below illustrates the general order for filling orbitals:

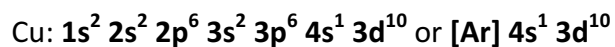
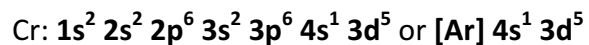
1s	→			
2s	→			2p
3s	→			3p
4s	→	3d	→	4p

Write the electron configuration for vanadium, using the noble gas shortcut.

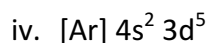
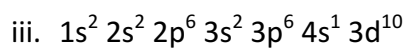
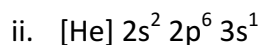
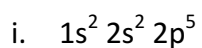
1. The electron configuration of an atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$, which refers to the ground state of
 - A. Mn.
 - B. Fe.
 - C. Ca.
 - D. Ti.

Exceptions to the Aufbau Principle

Certain metals will violate the aufbau principle because they have increased stability due to a half-filled or filled d sublevel. The electron configurations of chromium and copper, two exceptions to the aufbau principle, are shown below:



2. Use the periodic table to give the symbol of the element for each of the following electron configurations:

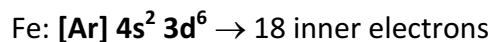


11.04

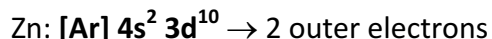
Categories of Electrons

There are three categories of electrons in an atom:

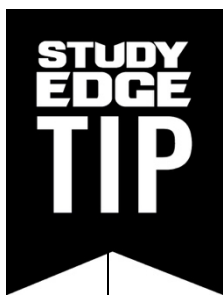
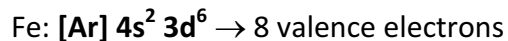
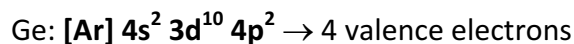
- **Inner (core) electrons** – The electrons in the lower energy levels of an atom up to the most immediate previous noble gas, plus any completed transition series



- **Outer electrons** – The electrons in the highest energy level (n value) of an atom



- **Valence electrons** – The electrons involved in compound formation
For main group elements (1A–8A), the valence electrons *are* the outer electrons. For the transition metals, all the $(n - 1)d$ electrons are counted as part of the valence electrons unless the d sublevel is filled.



Valence electrons = Total electrons – Inner electrons

1. Select the false statement below.

- A. The V atom has 2 outer electrons.
- B. The P atom has 5 valence electrons.
- C. The Cu atom has 18 inner electrons.
- D. The Na atom has 10 inner electrons and 1 outer electron.

11.05

Trends in Atomic Radii

Atomic radius – When two atoms of the same element are joined together, one half of the distance between the nuclei of the two atoms is the atomic radius.

- The atomic radius is the best way to predict the size of an atom.
- There are two trends that help predict the size of an atom:
 - As the atomic number increases down a group, the charge on the nucleus increases and the number of occupied energy levels increases.
 - There is an increase in nuclear charge and an increase in shielding, but the shielding effect is greater than the effect of the increase in nuclear charge.
 - This results in an increase in size down a group.
 - Across a period, each element has one more proton and electron than the preceding element.
 - The shielding effect is constant for all elements in a period, while the nuclear charge pulls the electrons in the highest occupied energy level closer to the nucleus.
 - This results in an increase in size, right to left, across a period.
- In general, atomic radii increase from top to bottom within a group and increase from right to left across a period.

1. In each pair, which element is larger?

i. Na, Rb

ii. Cl, Na

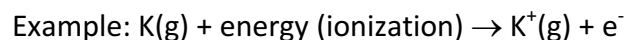
iii. Se, Br

11.06

Trends in Ionization Energy

Ionization energy – The energy required to remove an electron from an atom

- Ionization energy is measured when the atom is in its gaseous state.



- The first ionization energy is the energy required to remove the first electron from an atom.
 - In general, the first ionization energy trend is the *opposite* of the atomic radius trend.
 - First ionization energy tends to increase from bottom to top within a group and increase from left to right across a period.
 - As the atomic number decreases from bottom to top in a group, the valence electrons are closer to the nucleus, so they are harder to remove.
 - As you go left to right across a period, the increase in nuclear charge causes an increase in the attraction between the nucleus and an electron.
1. Using only the periodic table, rank the elements in each of the following sets in order of *decreasing* first ionization energy:
- i. P, Mg, O
 - ii. K, Ca, Sr
 - iii. I, Xe, Rb

11.07

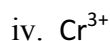
Electron Configuration for Ions

- An atom is electrically neutral because it has an equal number of protons and electrons. An ion forms when an atom or group of atoms loses or gains electrons.
 - A **cation**, or a positively charged ion, is produced when an atom loses one or more valence electrons.
 - An **anion**, or a negatively charged ion, is produced when an atom gains one or more valence electrons.
- The first electron removed by an atom will come from the highest occupied energy level (n) because those electrons experience less nuclear charge.

If there is more than one subshell in the highest occupied energy level, the atom removes electrons in the f subshell first, then the d subshell, then the p subshell, and finally the s subshell.
- The first electron gained by an atom will go to the lowest unoccupied energy level that has room.

If there is more than one subshell in the lowest unoccupied energy level, the atom gains electrons in the s subshell first, then the p subshell, then the d subshell, and finally the f subshell.
- In becoming an ion, an atom gains or loses electrons to reach a noble gas configuration. Noble gases have very low reactivity because their highest occupied energy level is full of electrons, which makes them very stable.
- Some cations do not form cations with a noble gas configuration; instead, they become cations with two different stable configurations:
 - **Pseudo-noble gas configuration** occurs when an element empties out its highest s and p subshells to leave a filled d subshell behind.
 - **Inert pair configuration** occurs when an element empties out its highest occupied p subshell to leave filled s and d subshells.

1. Write the electron configuration for the following ions.



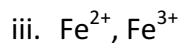
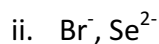
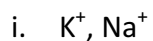
2. How many unpaired electrons are present in the full ground-state electron configuration of the monatomic ion most likely to be formed by Mg?
- A. 1
 - B. 2
 - C. 3
 - D. No unpaired electrons

11.08

Trends in Ionic Radii

- The **ionic radius** is the approximate distance from the center of an ion to its highest occupied electron orbital. We observe the following phenomena:
 - *Cations are smaller than the parent atom* – When electrons are removed, the nuclear charge increases, so the size of the atom decreases.
Example: $\text{Na} > \text{Na}^+$
 - *Anions are larger than the parent atom* – When electrons are gained, the nuclear charge decreases, so the size of the atom increases.
Example: $\text{Cl}^- > \text{Cl}$
- Ionic radii follow the trend of atomic radii, increasing top to bottom within a group and increasing right to left across a period.

1. In each pair, which ion is smaller?



PERIODIC TABLE OF THE ELEMENTS

1A 1	2A 2											3A 13	4A 14	5A 15	6A 16	7A 17	8A 18 2								
1 H 1.008	2 He 4.003											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18								
3 Li 6.941	4 Be 9.012											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95								
11 Na 22.99	12 Mg 24.31	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9 10	11 11	12 12	13 13	14 14	15 15	16 16	17 17	18 18									
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80								
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (99)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3								
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)								
87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Rf (261)	105 Ha (262)	106 106	107 107	108 108	109 109																	
												66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0								
												98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)								
												65 Tb 158.9	64 Gd 157.2	63 Eu 152.0	62 Sm 150.4	61 Pm (145)	60 Nd 144.2	59 Pr 140.9	58 Ce 140.1						
												97 Bk (247)	96 Cm (247)	95 Am (243)	94 Pu (244)	93 Np 237.0	92 U 238.0	91 Pa 231.0	90 Th 232.0						
												Lanthanides							Actinides						