

Oakland Schools Chemistry Resource Unit

Electrons & Quantum Mechanics

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Electrons and Quantum Mechanics

Content Statements:

C4.8x Electron Configuration

Electrons are arranged in main energy levels with sublevels that specify particular shapes and geometry. Orbitals represent regions of space in which an electron may be found with a high level of probability. Each defined orbital can hold two electrons, each with a specific spin orientation. The specific assignment of an electron to an orbital is determined by a set of 4 quantum numbers. Each element and therefore, each position in the periodic table is defined by a unique set of quantum numbers.

C2.4x Electron Movement

For each element, the arrangement of electrons surrounding the nucleus is unique. These electrons are found in different energy levels and can only move from a lower energy level (closer to nucleus) to a higher energy level by absorbing energy in discrete packets. The energy content of the packets is directly proportional to the frequency of the radiation. These electron transitions will produce unique absorption spectra for each element. When the electron returns from an excited (high energy state) to a lower energy state, energy is emitted in only certain wavelengths of light, producing an emission spectra.

Content Expectations:

C4.8i Describe the fact that the electron location cannot be exactly determined at any given time.

C48e Write the complete electron configuration of elements in the first four rows of the periodic table.

C4.8f Write kernel structures for main group elements.

C4.8g Predict oxidation states and bonding capacity for main group elements using their electron structure.

C4.8h Describe the shape and orientation of s and p orbitals

C2.4a Describe energy changes in flame tests of common elements in terms of the (characteristic) electron transitions.

C2.4b Contrast the mechanism of energy changes and the appearance of absorption and emission spectra.

C2.4c Explain why an atom can absorb only certain wavelengths of light.

C2.4d Compare various wavelengths of light (visible and non visible) in terms of frequency and relative energy.

Instructional Background Information:

Nature of electrons

Although the Bohr model of the atom makes a nice picture (in fact, it is the one normally used to represent an atom), it is no longer considered a good description of the behavior of electrons. There are two major reasons: (1) it only works for one-electron atoms and (2) there is no reason given for electrons to behave in that manner. Instead, the current theory to explain electrons is quantum mechanics.

(<http://www.wwnorton.com/college/chemistry/gilbert/overview/ch3.htm>)

The modern model of the atom does not tell an exact location of an electron. In order to understand an electron, it is important to have a basic knowledge of waves. According to wave mechanics electrons do not move about the atom in a definite pattern. An electron orbit in a quantum-mechanical model does not look like a planetary orbital. The orbit in this sense has no direct physical meaning. However, we are able to propose statistically where an electron is found or where it spends the majority of its time within an atom. This is called the probability density which is found by squaring the wave function (the electrons matter-wave position in three dimensions). For a given energy level it can be depicted by a probability density diagram or an electron cloud. The cloud is an imaginary picture of the electron changing position, not the image of an electron diffusing a cloud of charge.

The probable location of an electron depends on the amount of energy an electron has. The electrons with a low amount of energy are found in the lowest energy levels (closest to the nucleus); electrons with a high amount of energy are found on the outermost energy levels (furthest from the nucleus).

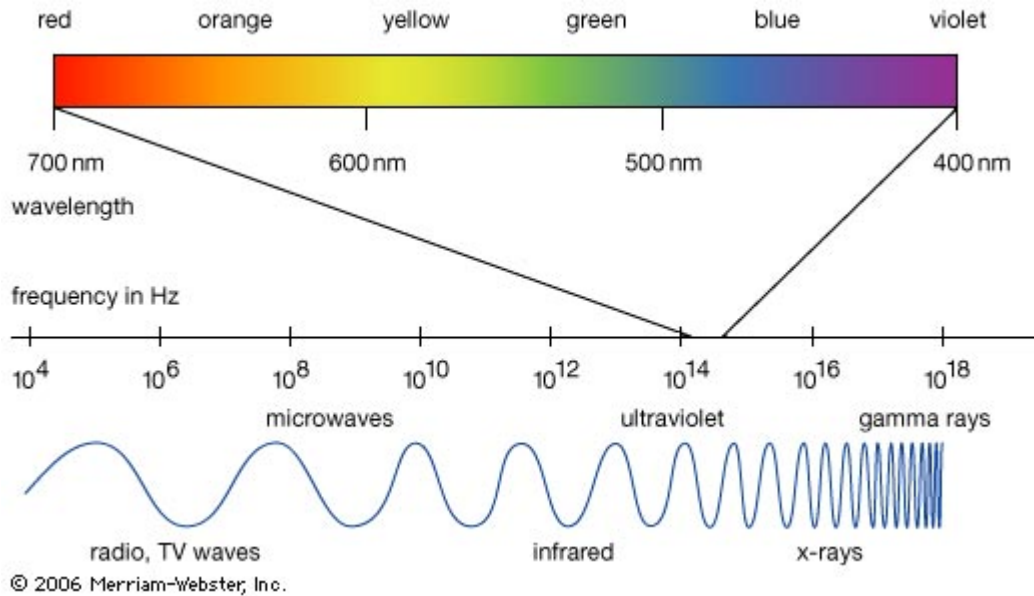
<http://education.jlab.org/jsat/powerpoint/13>

Spectroscopy

Gamma rays, X-rays, ultra violet rays, infrared, microwaves radio waves and visible light are several different types of electromagnetic (EM) radiation (or waves). All types of electromagnetic radiation are produced by alternating electrical and magnetic fields (or electromagnetic fields). Understanding the wave model helps explain characteristics of both waves and particles. Real world applications include understanding of how rainbows form, why objects look distorted underwater and how magnifying glasses work. All electromagnetic waves have the same speed, 3.00×10^8 m/s in a vacuum. Other types of media such as air or water only slightly slow the speed of light and other types of EM waves. When a wave strikes the boundary of the media the change in speed causes a change in direction, causing the wave to continue at a different angle. This process causes white light to disperse or separate into its component colors, similar to how light looks when it passes through a prism.

EM waves vary in types according to changes in both frequency (ν) and wavelength (λ). Wavelengths are measure in nm (nanometers), and measure the distance between any point on a wave and the corresponding point on the next crest (or trough) of the wave

or the distance that the wave travels during one cycle. Frequency which is the number of cycles that a wave undergoes per second is measured in Hz (1/s (seconds)). Amplitude is the height of the crest or depth of the trough of each wave and is related to the intensity of the radiation, which is perceived as brightness in the case of visible light.

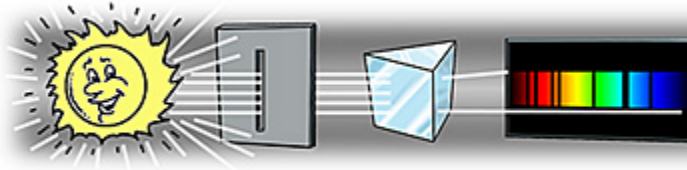


Waves transmit energy. If you look at an object bobbing in the water, the object does not move along the surface of the wave but bobs up and down with the wave as it passes. The wave does not carry the water; it transmits energy through the water. The amount of energy is determined by the frequency and amplitude. Energy of waves provides evidence that describes the current theory of electrons in atoms.

Emission and Absorption

Spectra <http://www.colorado.edu/physics/2000/quantumzone/fraunhofer.html>

When sunlight is sent through a thin slit and then through one of the prisms, it formed a rainbow-colored spectrum but the spectrum also contains a series of dark lines.

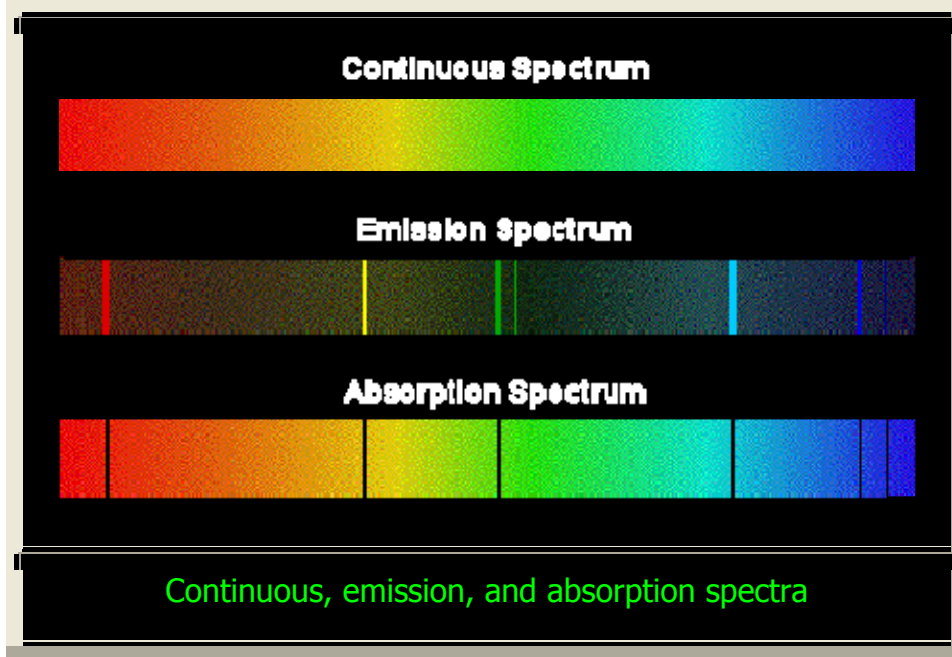


This also happens when an element is *heated*. In terms of the Bohr model, heating the atoms gives them some extra energy, so some of their electrons can jump up to higher energy levels. However electrons are not able to stay in the higher levels. When one of the electrons drops back down to a lower level, it emits a photon --at one of that element's special frequencies. This is called an emission spectrum. But there is another way in which elements can produce spectra. Suppose that instead of a heated sample of some element, you have the element in the form of a relatively cool gas. Now let's say that a source of white light-- containing all visible wavelengths--is shining behind the gas. When photons from the light source make their way through this gas, some of them can interact with the atoms--provided that they have just the right frequency to bump an electron of that element up to a higher energy level. Photons at those particular frequencies are thus absorbed by the gas. However, the atoms are "transparent" to photons of other frequencies, so all those other frequencies would come through okay. Then the spectrum of light that had been through the gas would just have some gaps in it, at the frequencies that were absorbed.



The spectrum with these missing frequencies is called an absorption spectrum. (Note that the dark lines in an absorption spectrum appear at exactly the same frequencies as the bright lines in the corresponding emission spectrum.)

Under very careful examination, the "continuous" spectrum of sunlight turns out to be an absorption spectrum. In order to reach earth, sunlight needs to pass through the sun's atmosphere, which is a lot cooler than the part of the sun where light is emitted. Gases in the atmosphere thus absorb certain frequencies, creating the 600 or so dark lines. (These are called Fraunhofer lines; he was the first to submit these observations.)



<http://csep10.phys.utk.edu/astr162/lect/light/absorption.html>

Emission spectra are produced by thin gases in which the atoms do not experience many collisions (because of the low density). The emission lines correspond to photons of discrete energies that are emitted when excited atomic states in the gas make transitions back to lower-lying levels.

A **continuum spectrum** results when the gas pressures are higher. Generally, solids, liquids, or dense gases emit light at all wavelengths when heated.

An **absorption spectrum** occurs when light passes through a cold, dilute gas and atoms in the gas absorb at characteristic frequencies; since the re-emitted light is unlikely to be emitted in the same direction as the absorbed photon, this gives rise to dark lines (absence of light) in the spectrum.

<http://jersey.uoregon.edu/vlab/elements/Elements.html>

Summary

- An electron absorbs heat or electrical energy and is promoted to a higher level
- The electron returns to the original level and emits the difference as a specific electromagnetic radiation.
- The wavelength seen is related to the energy of the emission by Planck's equation $E=h\nu$

E = energy of the emission

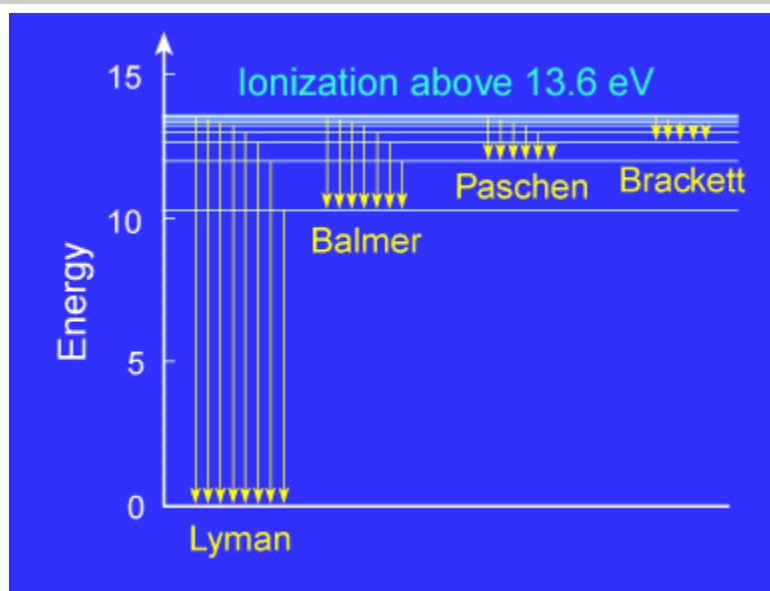
h = Planck's constant (6.02×10^{-34})

ν = frequency of the radiation (the frequency is related to the wavelength by $c = \lambda\nu$, c is the speed of light and λ is the wavelength)

Example:-

Hydrogen spectrum transitions

giving series of lines in the visible range due to transitions --> $n=2$:(Balmer series)



http://ibchem.com/IB/ibnotes/full/ato_hm/12.2.htm - Periodic table showing absorption and emission spectrum.

<http://www.wwnorton.com/college/chemistry/gilbert/overview/ch3.htm#spectrum> - Light emission and absorption Tutorial for Na.

Electron Configuration

Atoms are composed of three smaller parts—protons, neutrons, and electrons. Protons are positively charged, neutrons have no charge, and electrons are negatively charged. A neutral atom contains an equal number of protons and electrons. The protons and neutrons together form the nucleus of the atom, occupying less than 0.01% of the total atomic volume but accounting for 99.99% of the atomic mass. Electrons, although contributing negligibly to the mass, spend their time in the remaining volume.

This means that most of the volume of an atom is essentially “nothingness.” The number of protons in the nucleus of an atom determines the element, but the electrons determine the element’s chemical behavior. The outermost or valence electrons are those that can be shared with other atoms to form molecules, or can be transferred to or from other atoms, forming ionic salts. To understand how valence electrons determine the chemistry of an element, the internal structure of the atom must be examined in greater detail.

Electrons occupy specific “spaces” or energy levels in an atom. These energy levels are quantized; the location of electrons in terms of energies (the electron configuration) determines the nature and number of chemical bonds that each element forms. The chemical and physical properties of an element are determined by its electron configuration. These properties are the direct consequence of electron configuration, are observed in such phenomena as fireworks, neon lights, television, and colorful fireplace logs. In each instance, the distinctive colors are due to excitation of valence electrons.

More about electron configurations

Note:

It should be remembered that in spite of this order the 4s electrons fill up before the 3d electrons when applying the Aufbau principle, and that there are anomalies in the configurations of chromium, $[\text{Ar}] 4s^1 3d^5$ and copper, $[\text{Ar}] 4s^1 3d^{10}$ due to the extra stability of half full and full sets of “d” orbitals respectively. The number of orbitals at each energy level

This is usually shown in graphical form. Click on the link to see the [Aufbau principle in practice](#).

The number of orbitals at each level : $s=1$, $p=3$, $d=5$, $f=7$

Level 1: has only one s orbital

Level 2: has one s and three p orbitals

Level 3: has one s, three p and five d orbitals

Level 4: has one s, three p, five d and seven f orbitals

State the Aufbau principle. Reference should be made to Hund's rule

The Aufbau principle just means the way the electrons fit into the atomic orbitals in order of ascending energy. The first electron goes into the lowest energy orbital available (the 1s orbital) the next electron pairs up with it in the same orbital and the third electron (that of lithium) fits into the next orbital up, the 2s orbital.

The rules for filling up the orbitals are as follows:

1. Electrons always enter the orbital of lowest energy
2. If there are two or more degenerate orbitals (meaning that they have the same energy) then the electrons will singly occupy the degenerate orbitals until the orbitals are all singly occupied, after which they will pair up one at a time. This is known as Hund's rule
3. There cannot be more than two electrons in any one orbital.

The Aufbau (building up) Principle:

1. The number of electrons in an atom is equal to the atomic number;
2. Each added electron will enter the orbitals in the order of increasing energy;
3. An orbital cannot take more than 2 electrons.

<http://www.chemtutor.com/struct.htm#con>

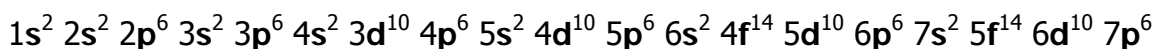
The electrons are in discrete pathways or shells around the nucleus. There is a ranking or hierarchy of the shells, usually with the shells further from the nucleus having a higher energy. As we consider the electron configuration of atoms, we will be describing the **ground state** position of the electrons. When electrons have higher energy, they may move up away from the nucleus into higher energy shells. As we consider the electron configuration, we will be describing the ground state positions of the electrons.

A hydrogen atom has only one proton and one electron. The electron of a hydrogen atom travels around the proton nucleus in a shell of a spherical shape. The two electrons of helium, element number two, are in the same spherical shape around the nucleus. The first shell only has one subshell, and that subshell has only one orbital, or pathway for electrons. Each orbital has a place for two electrons. The spherical shape of the lone orbital in the first energy level has given it the name 's' orbital. Helium is the last element in the first period. Being an inert element, it indicates that that shell is full. Shell number one has only one s subshell and all s subshells have only one orbital. Each orbital only has room for two electrons. So the first shell, called the K shell, has only two electrons.

Beginning with lithium, the electrons do not have room in the first shell or energy level. Lithium has two electrons in the first shell and one electron in the next shell. The first shell fills first and the others more or less in order as the element size increases up the Periodic Chart, but the sequence is not immediately obvious. The second energy level has room for eight electrons. The second energy level has not only an **s** orbital, but also a **p** subshell with three orbitals. The **p** subshell can contain six electrons. The **p** subshell has a shape of three dumbbells at ninety degrees to each other, each dumbbell shape being one orbital. With the **s** and **p** subshells the second shell, the L shell, can hold a total of eight electrons. You can see this on the periodic chart. Lithium has one electron in the outside shell, the L shell. Beryllium has two electrons in the outside shell. The **s** subshell fills first, so all other electrons adding to this shell go into the **p** subshell. Boron has three outside electrons, carbon has four, nitrogen has five, oxygen has six, and fluorine has seven. Neon has a full shell of eight electrons in the outside shell, the L shell, meaning the neon is an inert element, the end of the period.

Beginning again at sodium with one electron in the outside shell, the M shell fills its **s** and **p** subshells with eight electrons. Argon, element eighteen, has two electrons in the K shell, eight in the L shell, and eight in the M shell. The fourth period begins again with potassium and calcium, but there is a difference here. After the addition of the **4s** electrons and before the addition of the **4p** electrons, the sequence goes back to the third energy level to insert electrons in a **d** shell.

The shells or energy levels are numbered or lettered, beginning with K. So K is one, L is two, M is three, N is four, O is five, P is six, and Q is seven. As the **s** shells can only have two electrons and the **p** shells can only have six electrons, the **d** shells can have only ten electrons and the **f** shells can have only fourteen electrons. The sequence of addition of the electrons as the atomic number increases is as follows with the first number being the shell number, the **s**, **p**, **d**, or **f** being the type of subshell, and the last number being the number of electrons in the subshell.



It is tempting to put an **8s²** at the end of the sequence, but we have no evidence of an R shell. One way to know this sequence is to memorize it. There is a bit of a pattern in it. The next way to know this sequence is to SEE IT ON THE PERIODIC CHART. As you go from hydrogen down the chart, the Groups 1 and 2 represent the filling of an **s** subshell. The filling of a **p** subshell is shown in Groups 3 through 8. The filling of a **d** subshell is represented by the transition elements (ten elements), and the filling of an **f** subshell is shown in the lanthanide and actinide series (fourteen elements).

Here is a copy of the periodic chart as you have usually seen it.

H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	Lu 71	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Lr 103	Db 104	Jl 105	Rf 106	Bh 107	Hn 108	Mt 109									

La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70
Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102

Here is the same chart re-arranged with the Lanthanides and Actinides in their right place and Group I and II afterward. Both of these charts are color coded so that the elements with the 2s subshell on the outside (H and He) are turquoise. All other elements with an s subshell on the outside (Groups I and II) are outlined in blue. Lanthanides and actinides are in grey. Other transition elements are in yellow, and all of the elements that have a p subshell as the last one on the outside are in salmon color.

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

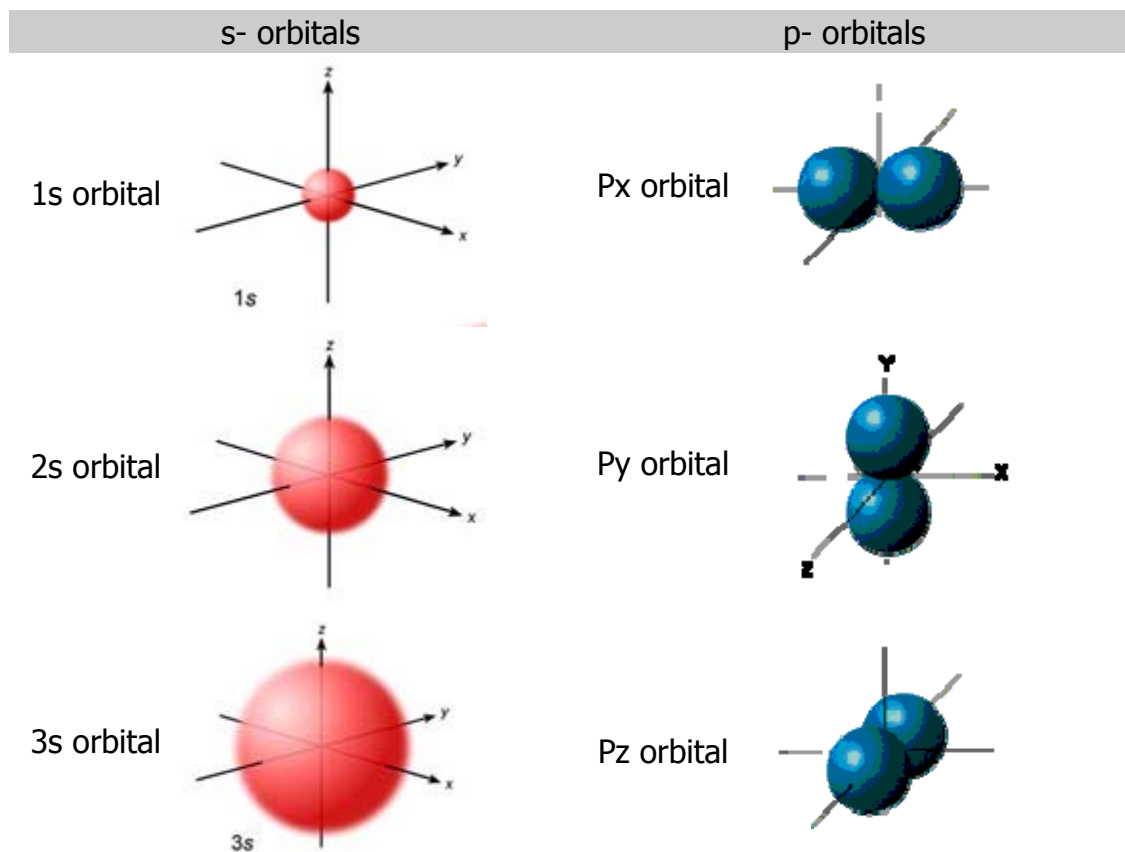
You may be able to see it better with the subshell areas labeled.

																		1 s	H 1																
																1 s	He 2	2 s	Li 3	Be 4															
										2 p	B 5	C 6	N 7	O 8	F 9	Ne 10	3 s	Na 11	Mg 12																
										3 p	Al 13	Si 14	P 15	S 16	Cl 17	Ar 18	4 s	K 19	Ca 20																
						3 d	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	4 p	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36	5 s	Rb 37	Sr 38									
						4 d	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	5 p	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54	6 s	Cs 55	Ba 56									
4 f	La 57	Ce 58	Pr 59	Nd 60	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	5 d	Lu 71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	6 p	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn	7 s	87 Fr	88 Ra
5 f	Ac 89	Th 90	Pa 91	U 92	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	6 d	103 Lr	104 Db	105 Jl	106 Rf	107 Bh	108 Hn	109 Mt				7 p	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og	8 s	119 Uu	120 Uub

There are several other schemes to help you remember the sequence.

Shapes of orbitals

The shape of the **s** subshells is spherical around the nucleus. The shape of the **p** subshells is the shape of three barbells at ninety degrees to each other around the nucleus. The shape of the **d** and **f** subshells is very complex.



Electron configuration is the "shape" of the electrons around an atom, that is, which energy level (shell) and what kind of orbital it is in. The shells were historically named for the chemists who found and calculated the existence of the first (inner) shells. Their names began with "K" for the first shell, then "L," then "M," so subsequent energy levels were continued up the alphabet. The numbers one through seven have since been substituted for the letters. Notice that I have included an "R" shell (#8) that is purely fantasy but makes the chart symmetrical.

The electron configuration is written out with the first (large) number as the shell number. The letter is the orbital type (either **s**, **p**, **d**, or **f**). The smaller superscript number is the number of electrons in that orbital.

Use this scheme as follows. You first must know the orbitals. An **s** orbital only has 2 electrons. A **p** orbital has six electrons. A **d** orbital has 10 electrons. An **f** orbital has 14 electrons. You can tell what type of orbital it is by the number on the chart. The only

exception to that is that "8" on the chart is "2" plus "6," that is, an s and a p orbital. The chart reads from left-to-right and then down to the next line, just as English writing. Any element with over 20 electrons in the electrically neutral unattached atom will have all the electrons in the first row on the chart. For instance, scandium, element #21, will have all the electrons in the first row and one from the second. The electron configuration of scandium is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ Notice that the $2s^2 2p^6$ and $3s^2 3p^6$ came from the eights on the chart (2+6). Notice that the other electron must be taken from the next spot on the chart and that the next spot is the first spot on the left in the next row. It is a 3d spot due to the "10" there and only one more electron is needed, hence $3d^1$.

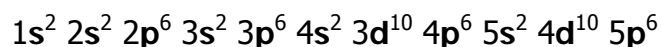
The totals on the right indicate using whole rows. If an element has an atomic number over thirty-eight, take all the first two rows and whatever more from the third row. Iodine is number fifty-three. For its electron configuration you would use all the electrons in the first two rows and fifteen more electrons. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$ from the first two rows and $4d^{10} 5p^5$ from the third row. You can add up the totals for each shell at the bottom. Full shells would give you the totals on the bottom.

We have included an R shell (#8) even though there is no such thing yet proven to exist. The chart appears more symmetrical with that shell included. The two electrons from the R shell are in parentheses. We have not yet even made elements that have electrons in the p subshell of the Q shell.

ELECTRON CONFIGURATION CHART

K	L	M	N	O	P	Q	R	
1	2	3	4	5	6	7	8	
s	sp	spd	spdf	spdf	spd	sp	s	
2	8	8	2					20
		10	6	2				38
			10	6	2			56
			14	10	6	2		88
				14	10	6	2	
-----		-----		-----		-----		
2	8	18	32	32	18	8	2	TOTALS

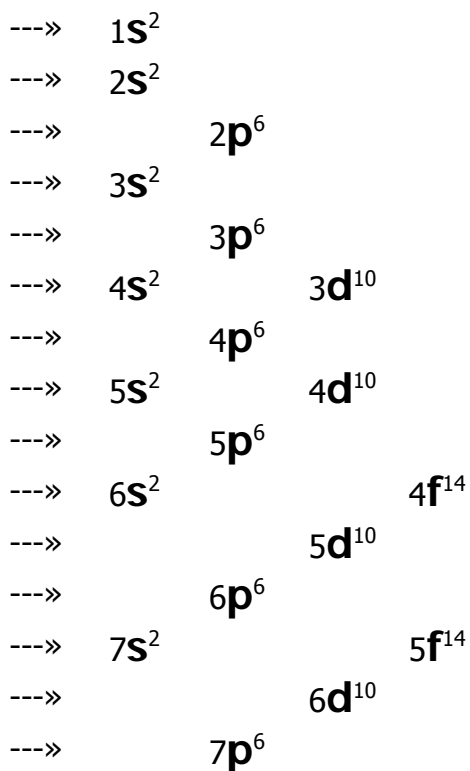
Here is another way to consider the same scheme. The inert elements appear at the end of either the first two, an eight, a six. Wherever there is the six of a p subshell there is the two of an s subshell above it to make eight electrons in the outer full shell of a noble gas. The electron configuration for xenon is:



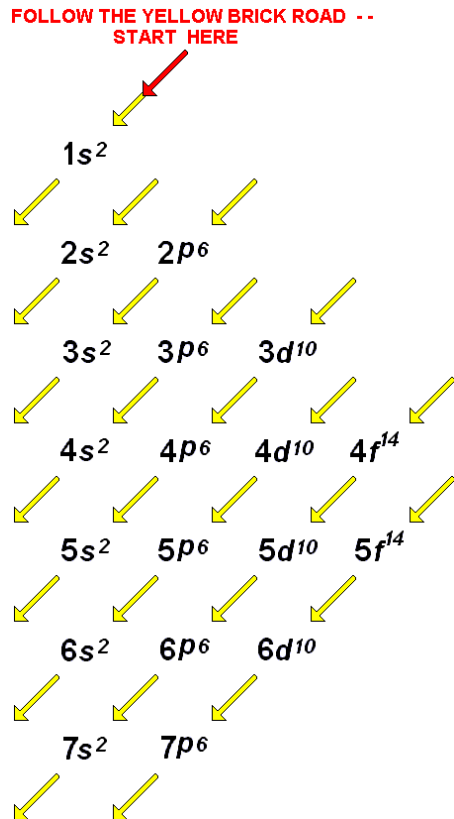
“Und” is the undiscovered inert element that would be below radon on the periodic chart.

Another type of electron configuration chart is below. These are more commonly known schemes. All you have to do is follow the arrows through the points to find the sequence. Add up the number of electrons as you go, and stop when you have equaled or almost exceeded the number. There have been a large number of variations on this idea, but they all work the same. Arrange the subshells in a slanted order and go through the array in straight lines, as in the first scheme, or arrange the subshells in a straight line and go through the array in slanted lines, as in the second scheme. In these schemes the inert elements appear after the first **s** subshell and after every **p** subshell. As the other type, this scheme type has its advantages and disadvantages, but they all lead to the same sequence.

COMMON ELECTRON CONFIGURATION SCHEME A



COMMON ELECTRON CONFIGURATION SCHEME B



Any of these schemes, if used correctly, will give you the same thing, the sequence of the addition of the electrons to the shells. This pattern is correct for all of the elements that are not Transitional Elements or Lanthanides or Actinides. Of the Transitional Elements and Lanthanides and Actinides about one third of the elements do not follow the pattern. The Periodic Chart below is arranged sideways to show the electron configuration by shell. As you work with the schemes for finding the electron configuration of elements, you can check to see if your answer is correct by adding the electrons in each shell (downwards in the first scheme) and comparing with the Sideways Periodic Chart. The elements that do not fit the pattern have an asterisk by them. In the Transition Elements that do not follow the scheme, only the **s** subshell of the outer shell and the **d** subshell of the next to last shell have some trading between them. In the Lanthanide and Actinide series any trading of electrons are between the **d** subshell of the next to last shell and the **f** subshell of the second to last shell, the one filling as the elements progress up that series.

The octet rule as seen on the periodic table

The **octet rule** states that atoms are most stable when they have a full shell of electrons in the outside electron ring. The first shell has only two electrons in a single **s** subshell. Helium has a full shell, so it is stable, an inert element. Hydrogen, though, has only one electron. It can lose an electron to become H^+ , a hydrogen ion or it can gain an electron to become H^- , a hydride ion. All the other shells have an **s** and a **p** subshell, giving them at least eight electrons on the outside. The **s** and **p** subshells often are the only valence electrons, thus the octet rule is named for the eight **s** and **p** electrons.

On the Periodic Chart with shell totals you can easily see the octet rule. A valence is a likely charge on an element ion. All of the Group 1 elements have one electron in the outside shell and they all have a valence of plus one. Group 1 elements will lose one and only one electron, so the single outside electron will become a single positive ion with a full electron shell of eight electrons (an octet) in the **s** and **p** subshells under it.

Group 2 elements all have two electrons in the outer shell and all have a valence of plus two. Beryllium can be a bit different about this, but all other Group 2 elements can lose two electrons to become +2 ions. They do not lose only one electron, but two or none.

The Transition Elements, Lanthanides, and Actinides are all metals. Many of them have varying valences because they can trade around electrons from the outer shell to the inner **d** or **f** subshells that are not filled. For this reason they sometimes appear to violate the octet rule.

Group 3 elements have a valence of plus three. Boron is a bit of an exception to this because it is so small it tends to bond covalently. Aluminum has a valence of +3, but some of the larger Group 3 elements have more than one valence.

The smallest Group 4 elements, carbon and silicon, are non-metals because the four electrons are difficult to lose the entire four electrons in the outer shell. Small Group 4 elements tend to make only covalent bonds, sharing electrons. Larger Group 4 elements have more than one valence, usually including +4.

Small Group 5 elements, nitrogen and phosphorus, are non-metals. They tend to either gain three electrons to make an octet or bond covalently. The larger Group 5 elements have more metallic character.

Small Group 6 elements, oxygen and sulfur, tend to either gain two electrons or bond covalently. The larger Group 6 elements have more metallic character.

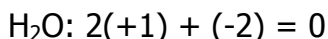
Group 7 elements all have seven electrons in the outer shell and either gain one electron to become a -1 ion or they make one covalent bond. The Group 7 elements are diatomic gases due to the strong tendency to bond to each other with a covalent bond.

All of the inert elements, the noble gases, have a full octet in the outside shell (or two in the first shell) and so do not naturally combine chemically with other elements.

Oxidation numbers

It is often useful to follow chemical reactions by looking at changes in the oxidation numbers of the atoms in each compound during the reaction. Oxidation numbers also play an important role in the systematic nomenclature of chemical compounds. By definition, the oxidation number of an atom is the charge that atom would have if the compound was composed of ions.

1. The oxidation number of an atom is zero in a neutral substance that contains atoms of only one element. Thus, the atoms in O_2 , O_3 , P_4 , S_8 , and aluminum metal all have an oxidation number of 0.
2. The oxidation number of simple ions is equal to the charge on the ion. The oxidation number of sodium in the Na^+ ion is +1, for example, and the oxidation number of chlorine in the Cl^- ion is -1.
3. The oxidation number of hydrogen is +1 when it is combined with a *nonmetal* as in CH_4 , NH_3 , H_2O , and HCl .
4. The oxidation number of hydrogen is -1 when it is combined with a *metal* as in LiH , NaH , CaH_2 , and $LiAlH_4$.
5. The metals in Group IA form compounds (such as Li_3N and Na_2S) in which the metal atom has an oxidation number of +1.
6. The elements in Group IIA form compounds (such as Mg_3N_2 and $CaCO_3$) in which the metal atom has a +2 oxidation number.
7. Oxygen usually has an oxidation number of -2. Exceptions include molecules and polyatomic ions that contain O-O bonds, such as O_2 , O_3 , H_2O_2 , and the O_2^{2-} ion.
8. The elements in Group VIIA often form compounds (such as AlF_3 , HCl , and $ZnBr_2$) in which the nonmetal has a -1 oxidation number.
9. The sum of the oxidation numbers in a neutral compound is zero.



10. The sum of the oxidation numbers in a polyatomic ion is equal to the charge on the ion. The oxidation number of the sulfur atom in the SO_4^{2-} ion must be +6, for example, because the sum of the oxidation numbers of the atoms in this ion must equal -2.

$$\text{SO}_4^{2-}: (+6) + 4(-2) = -2$$

11. Elements toward the bottom left corner of the periodic table are more likely to have positive oxidation numbers than those toward the upper right corner of the table. Sulfur has a positive oxidation number in SO_2 , for example, because it is below oxygen in the periodic table.

$$\text{SO}_2: (+4) + 2(-2) = 0$$

Oxidation state is a number assigned to an element in a compound according to some rules. This number enables us to describe oxidation-reduction reactions, and balancing redox chemical reactions. You are learning the skill to assign oxidation states (or oxidation numbers) to a variety of compounds and ions.

When an oxidation number is assigned to the element, it does not imply that the element in the compound acquires this as a charge, but rather that it is a convenient number to use for balancing chemical reactions. The guidelines for assigning oxidation states (numbers) are given below:

1. The oxidation state of any element such as Fe, H_2 , O_2 , P_4 , S_8 is zero (0).
2. The oxidation state of oxygen in its compounds is -2, except for peroxides like H_2O_2 , and Na_2O_2 , in which the oxidation state for O is -1.
3. The oxidation state of hydrogen is +1 in its compounds, except for metal hydrides, such as NaH, LiH, etc., in which the oxidation state for H is -1.
4. The oxidation states of other elements are then assigned to make the algebraic sum of the oxidation states equal to the net charge on the molecule or ion.
5. The following elements usually have the same oxidation states in their compounds:
 - o +1 for alkali metals - Li, Na, K, Rb, Cs;
 - o +2 for alkaline earth metals - Be, Mg, Ca, Sr, Ba;
 - o -1 for halogens except when they form compounds with oxygen or one another;

These rules are wordy because we have to point out the special cases such as H_2O_2 and Na_2O_2 . Rule 3 deals with hydride. Other than these, you may simply remember the oxidation states for H and O are +1 and -2 respectively in a compound, and oxidation of other elements can be assigned by making the algebraic sum of the oxidation states equal to the net charge on the molecule or ion.

For your practice, we provide some examples below. Please study the following examples and derive the oxidation state for all elements. The oxidation numbers of the key element are given in case you need help.

Element	Oxidation state	Compound or ion
H	+1	H ⁺
Group 1	+1	H ₂ O
	0	H ₂
	-1	NaAlH ₄
Cl	-1	Cl ⁻
Group 7	0	Cl ₂
	+1	ClO ⁻
	+3	ClO ₂ ⁻
	+4	ClO ₂
	+5	ClO ₃ ⁻
	+7	ClO ₄ ⁻
N	-3	NH ₃
Group 5	-2	N ₂ H ₄
	-1	NH ₂ OH
	0	N ₂
	+1	N ₂ O
	+2	NO
	+3	NO ₂ ⁻
	+4	NO ₂
	+5	NO ₃ ⁻

Terms and Concepts

Absorbance spectrum	Atomic motion	Aufbau Principle
Bright line spectrum	Chemical bond	Electromagnetic field
Electromagnetic radiation	Electromagnetic spectra	Electromagnetic wave
Electron	Electron configuration	Emission spectra
Energy level	Excited state	Kernel
Ground state	Orbitals	Oxidation States
Probability	Quantum energy	Quantum numbers
Release of energy	Sublevel	Valence electrons
Wave amplitude	Wavelength	

Electrons & Quantum Mechanics

Activity #1 - Locating an "s" Electron in an Atom by Analogy

Questions to be investigated

How is an electron cloud formed?

What is the probability that an electron will be found in a particular region or space of an atom?

How is the radius of an atom determined?

Objectives

Students will get insight into the meaning of abstract concepts and terms (orbitals, electron density, charge cloud, probability graphs, *etc.*) associated with the wave-mechanical model of an atom.

Students will determine the distribution of impacts of marble drops around the bulls eye of a target.

Students will obtain and interpret probability information on the distribution of marble drops around a bulls-eye of a target.

Teacher Notes

Expected Student Background

Students should review wave-mechanical model of the atom.

Pre-Laboratory Discussion

The materials and procedures of the activity may be illustrated during a brief pre-laboratory discussion. During this discussion, explain that the activity's purpose is to develop "experimentally" a "charge cloud" diagram and probability graph. Point out the analogy to comparable mathematically derived electric charge-cloud diagrams and probability graphs found in some student texts.

NOTE: Student graphs should begin at a radius of 0, where the probability of finding an electron is 0.

Teacher-Student Interaction

1. Be sure the marbles are caught after their first bounce.
2. Be sure marbles are dropped from a consistent height (2 m).
3. Ask students whether they expect each group to get the same pattern (the patterns should be similar, but not identical).
4. Do students expect a marble to land *exactly* in the middle?

Materials

(For 24 students working in pairs)

24 Targets (see sample targets in *Appendix*)

12 Typewriter carbon papers

48 Paper clips

12 Marbles

Safety Concerns

There are no unusual safety requirements, except to insure that students with marble-drop responsibilities climb carefully onto stable, safe structures.

Source

http://dwb4.unl.edu/chem_source_pdf/ChemSource.html

Procedure

Data Analysis and Concept Development

1. Locate the 95 dots (95% of the 100) that are closest to the bulls-eye. Using a pencil, draw a smooth (not wavy) curve to enclose the region containing these dots. This region represents the two-dimensional "orbital" of your "marble electron" (analogous to the region of space where an electron might be observed 95% of the time in an experiment to locate it in three-dimensional space).
2. Count the dots in each concentric circular area surrounding the bulls-eye and record the number in the table. Multiply the number of dots in each area by the radius of the region given in the table. The product represents the dots times centimeters. Enter your calculations in the table.

Number of circular area increment (n)	Maximum radial distance of area increment from bullseye	Area of increment ^a	Number of dots in radial increment	Dots times area
0	0	0	0	0
1	R ₁ - 0.9 cm	2.54 cm ²		
2	R ₂ - 2.6 cm	18.7 cm ²		
3	R ₃ - 4.4 cm	42.1 cm ²		
4	R ₄ - 6.3 cm	82.6 cm ²		
5	R ₅ - 8.0 cm	119 cm ²		

^aArea of increment = $\pi R_{n+1}^2 - \pi R_n^2$

Radius of maximum probability (from graph) _____

Figure 2. Score card.

3. Prepare a graph by plotting the dots time's centimeters along the vertical axis and the maximum radial distance of the area increment from the bulls-eye along the horizontal axis. Note that radial distance extends from the center of the bulls-eye to the midpoint of the circular area. Start the graph line at 0.0 and allow it to approach the x axis asymptotically (without touching the axis).
4. Identify the radius of maximum probability from the graph (peak of the curve) and record it in the space provided in the table. This is the distance from the bulls-eye where you are most likely to find a spot (marble or electron) when attempting to locate it (drop a marble).
5. Compare your results to the radial probability curve for the s -electron (Figure 3). Sketch the s -electron plot on the graph paper containing your experimental curve. Use a colored pencil for the s -electron curve.

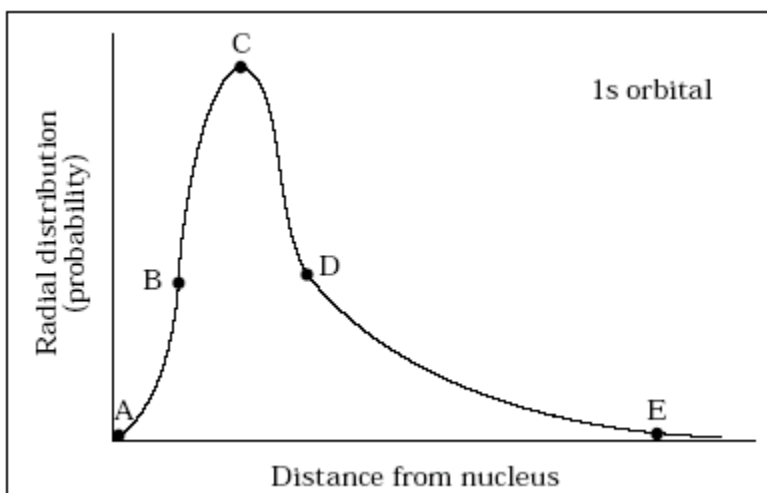


Figure 3. Radial probability curve for the s electron.

Assessment Ideas

- Why don't all the marbles land in the same spot? [There are many uncertainties that change the landing spot each time, height of hand, angle of release, etc.]
- Which point (A-E) is the greatest probability of finding the electron? [C]
 - Where is there zero probability of finding the electron? [A: the nucleus]
 - As the distance from the nucleus increases, what happens to the probability of finding an electron? [It peaks at some distance, C, then diminishes, but never reaches zero.]

Answers

Answers to Questions in Data Analysis and Concept Development

Number of circular area increment (n)	Maximum radial distance of area increment from bullseye	Area of increment*	Number of dots in radial increment	Dots times area
0	0	0	0	0
1	R ₁ = 0.9 cm	2.54 cm ²	1	2.54
2	R ₂ = 2.6 cm	18.7 cm ²	39	729
3	R ₃ = 4.4 cm	42.1 cm ²	36	1516
4	R ₄ = 6.3 cm	82.6 cm ²	19	1569
5	R ₅ = 8.0 cm	119 cm ²	5	595

$$* \text{Area of increment} = nR \frac{2}{n+1} - nR \frac{2}{n}$$

Radius of maximum probability (from graph) _____

Marble Drop Probability

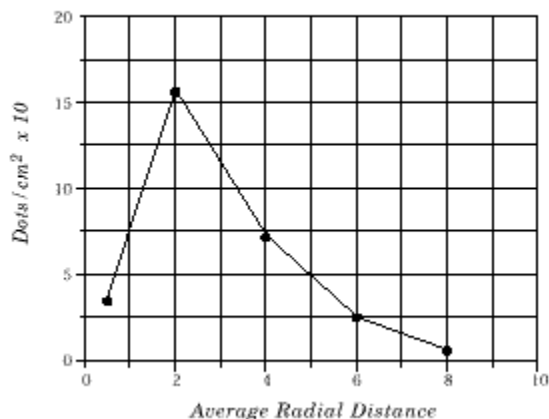


Figure 4. Sample data for Laboratory Activity 2.

Electrons & Quantum Mechanics

Activity 2: Locating an Electron in an Atom by Analogy

Introduction

If it were experimentally possible to determine how often an electron appeared at different points in space surrounding the nucleus, these data could be plotted on a three-dimensional graph. Boundaries could be established that would outline the regions of space within which the electron could be found 95% of the time. These regions of space, called orbitals, are commonly represented in texts as charge clouds, and physically through Styrofoam models.

Since it is impossible to conduct such an experiment, we shall perform one that, by analogy, will help give meaning to the concepts of probability distribution graphs, radii of maximum probability, maximum electron density, and orbitals.

In this laboratory activity, you will define regions of space around a bulls-eye in which there is a specific probability of locating a spot resulting from the impact of a marble dropped from a specified distance to the target. The spots represent points in space around the bulls-eye (analogous to the nucleus) where the marble (analogous to the electron) is observed in an experiment to locate the region of space within which the marble is most likely to strike (analogous to an atomic orbital).

You will also determine a radius of maximum probability (maximum spot density) by plotting the number of times the spots (analogous to places electrons might be observed) appear times the radius vs. distance from the bulls-eye (nucleus).

Purpose

1. To draw a physical analogy to some concepts associated with the wave-mechanical model of the atom.
2. To determine the distribution of impacts of marble drops around the bulls-eye of a target.
3. To obtain and interpret probability information on the distribution of marble drops around a bulls-eye of a target.

Procedure

1. Obtain two targets, a piece of typewriter carbon paper, and four paper clips.
2. Place the carbon paper between the two targets so that any impression made on the front target will be reproduced on the back target. Fasten with a paper clip at each corner.
3. With care, stand on a stool or laboratory bench with the marble in your hand. Extend your arm so it is about two meters above the floor. Have your partner place the target directly below your hand. Drop the marble on the target. Your partner must catch the marble on the first bounce (otherwise you'll have more than one mark for each drop of the marble). Repeat 99 times. A marble drop location will thus appear 100 times. The spots on the back target represent places where the marble was observed.

4. Return the front target, carbon paper, marble, and paper clips to locations designated by your teacher.

Data Analysis and Concept Development

1. Locate the 95 dots (95% of the 100) that are closest to the bulls-eye. Using a pencil, draw a smooth (not wavy) curve to enclose the region containing these dots. This region represents the two-dimensional orbital of your marble electron (analogous to the region of space where an electron might be observed 95% of the time in an experiment to locate it in three-dimensional space).

2. Count the dots in each concentric circular area surrounding the bulls-eye and record the number in the table. Multiply the number of dots in each area by the radius of the region given in the table. The product represents the dots times centimeters. Enter your calculations in the table.

*Area of increment = $\pi R^2_{n+1} - \pi R^2_n$

Number of circular area increment (n)	Maximum radial distance of area increment from bulls-eye	Area of increment*	Number of dots in radial increment	Dots times area
0	0	0	0	0
1	$R_1 = 0.9 \text{ cm}$	2.54 cm^2		
2	$R_2 = 2.6 \text{ cm}$	18.7 cm^2		
3	$R_3 = 4.4 \text{ cm}$	42.1 cm^2		
4	$R_4 = 6.3 \text{ cm}$	82.6 cm^2		
5	$R_5 = 8.0 \text{ cm}$	119 cm^2		

*Area of increment = $\pi R^2_{n+1} - \pi R^2_n$

Radius of maximum probability (from graph) _____

3. Prepare a graph by plotting the dots times centimeters along the vertical axis and the maximum radial distance of the area increment from the bulls-eye along the horizontal axis. Note that radial distance extends from the center of the bulls-eye to the midpoint of the circular area. Start the graph line at 0.0 and allow it to approach the x axis asymptotically (without touching the axis).

4. Identify the radius of maximum probability from the graph (peak of the curve) and record it in the space provided in the table. This is the distance from the bulls-eye where you are most likely to find a spot (marble or electron) when attempting to locate it (drop a marble).

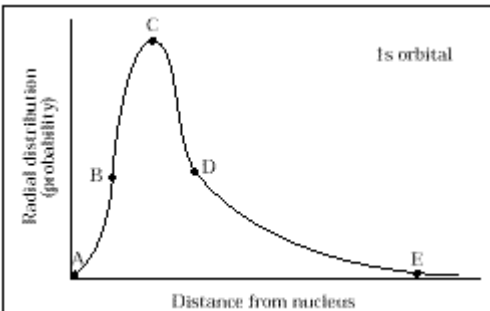


Figure 3 Radial probability curve for the 1s electron.

5. Compare your results to the radial probability curve for the s-electron (Figure 3). Sketch the s-electron plot on the graph paper containing your experimental curve. Use a colored pencil for the s-electron curve.

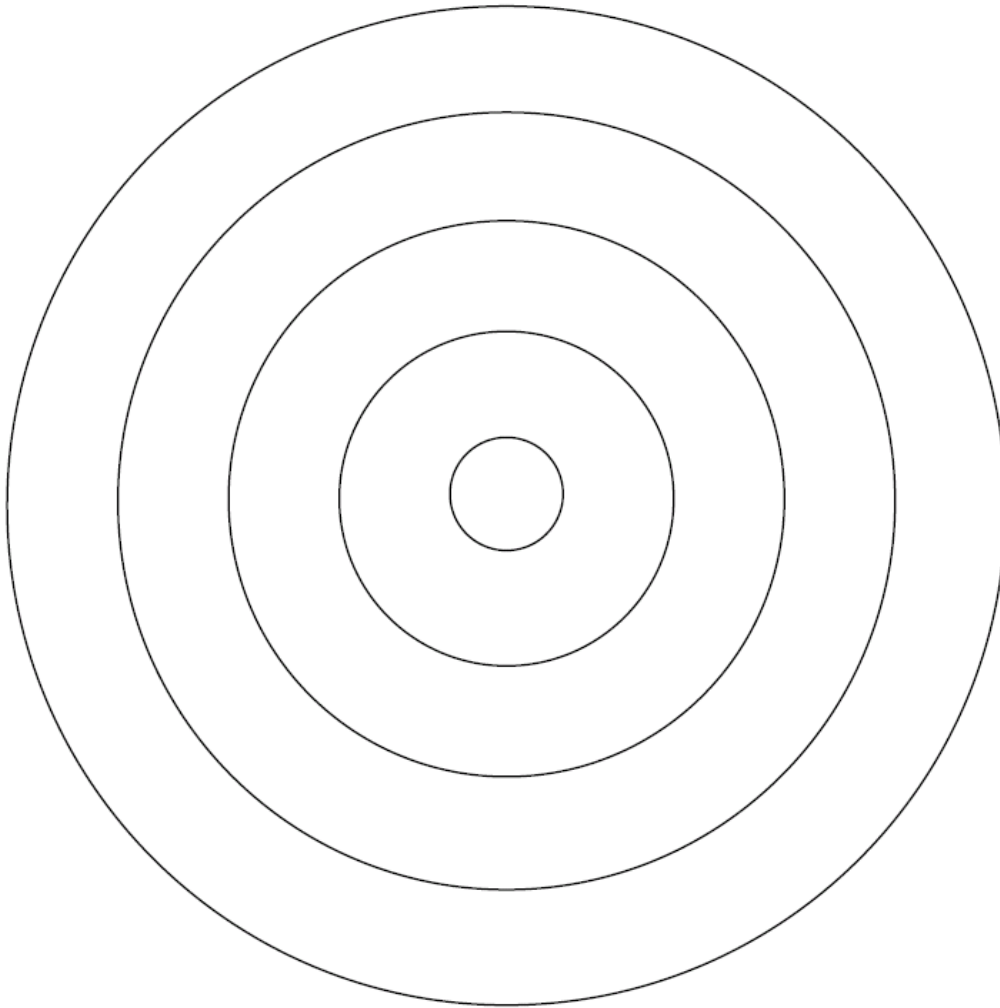
Turn in graph paper and sheet with carbon marks.



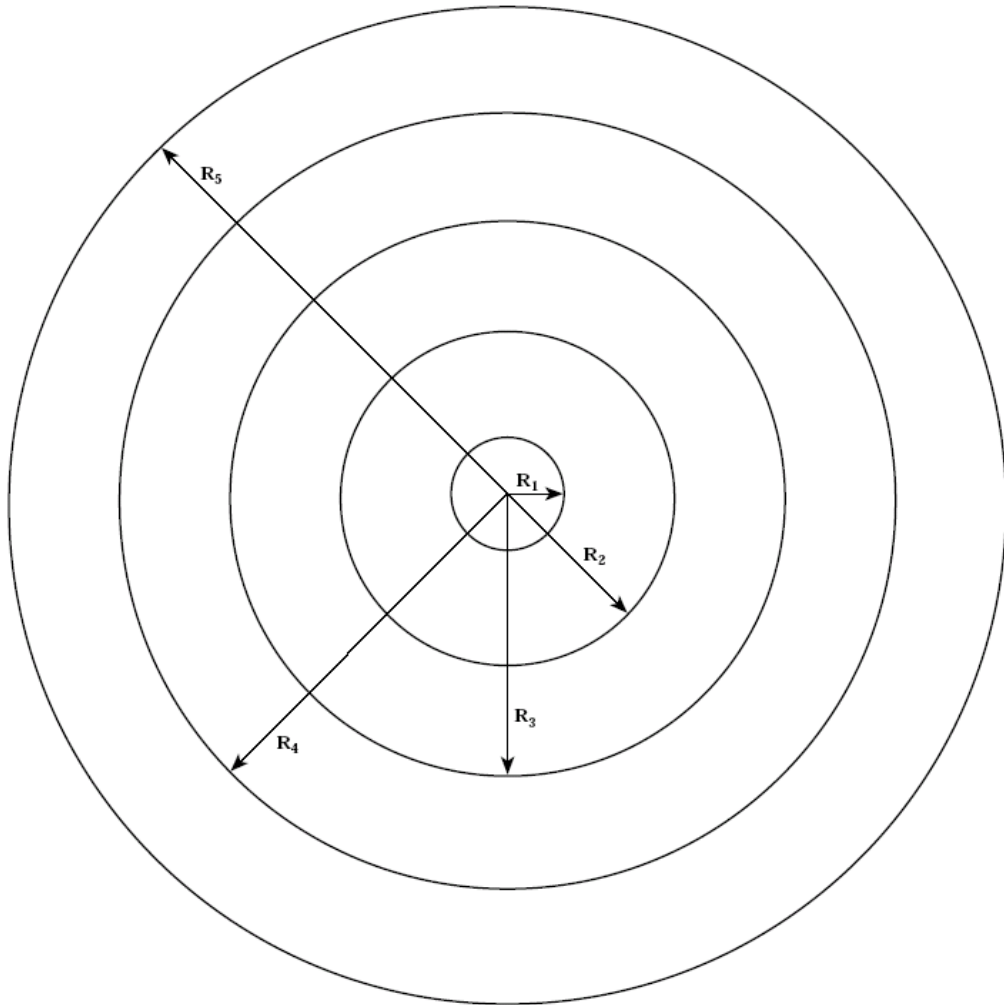
© 1992 by Sidney Harris



Target for Activity 2



Target for Activity 2 with Radii Shown



Electrons & Quantum Mechanics

Activity #2a - Excel Orbital Diagram

Questions to be investigated

How do electrons fall into the orbitals as they increase in number?

Objectives

Students will understand how electrons fill in order of increasing energy according to the Aufbau principle.

Teacher Notes

The set up for the program is fairly simple.

1. Open a new excel document.
2. Paste the following chart into the excel document.

Enter number of electrons here

36

4p	↑ ↓ ↑ ↓ ↑ ↓	31	34	32	35	33	36			
3d	↑ ↓ ↑ ↓ ↑ ↓ ↑ ↓ ↑ ↓	21	26	22	27	23	28	24	29	25 30
4s	↑ ↓	19	20							
3p	↑ ↓ ↑ ↓ ↑ ↓	13	16	14	17	15	18			
3s	↑ ↓	11	12							
2p	↑ ↓ ↑ ↓ ↑ ↓	5	8	6	9	7	10			
2s	↑ ↓	3	4							
1s	↑ ↓	1	2							

3. Highlight the yellow cells and change the text color to yellow (the arrows will become invisible)
4. Highlight the yellow cells. Select "Conditional Formatting" from the "Format" menu.
5. Pull down the first menu to change "Cell Value is" to "Formula is"
6. Copy: =IF(AN11<=\$A\$3,TRUE,FALSE) into the blank space to the right of "Formula is"

7. Click the "Format" button
8. Change the text color to black.
9. Click "OK"
10. Click "OK" (again)

Save the worksheet. The worksheet should now have black arrows (as seen above). As you change the numbers, you will get an electron turning black. This sheet will NOT address the Cr / Cu exceptions. You can use this in a computer lab or on a SmartBoard or TVator.

Materials

- Computers
- Periodic Table
- Excel

Procedure

See student document.

Assessment Ideas

Students can answer questions on the worksheet.

Name: _____ Date: _____
Hour: _____

Excel Electron Orbital Diagrams

Instructions:

1. Open the excel document provided by your teacher.
2. You will be changing the number of electrons in the top left corner. Start by typing the number "1" in that cell and pressing "Enter"
3. Increase the number to "2".
4. Repeat step 3 for numbers "3" through "36".

Questions:

1. What patterns did you notice? Describe *at least* two patterns that you noticed. Use *at least* 2 complete sentences to describe each pattern.
2. How many electrons fit in an "s" orbital?
3. How many electrons fit in a "p" orbital?
4. How many electrons fit in a "d" orbital?
5. Do you notice a pattern in the answers to questions 2 – 4? Describe it.
6. Using the pattern in question 5, how many electrons would you expect to fit in the next type of orbital (an "f" orbital)?

Electrons & Quantum Mechanics

Activity 2b-Instant Light - An Easy Chemiluminescence Demonstration

Introduction

Add several teaspoons of Instant Light crystals to water and watch as the solution instantly produces an eerie blue glow that will last for several minutes.

Chemical Concepts

- Chemiluminescence
- Oxidation–Reduction

Materials

Luminol, 0.2 g

2-3 Beakers 400-mL

Potassium ferricyanide, $K_3Fe(CN)_6$

4 g Magnetic stirrer and stir bar

64 g Clorox 2,[®] powder form, 64 g

or Instant Light Kit Distilled or deionized water, 400 mL

Safety Precautions

Potassium ferricyanide will emit poisonous fumes of hydrogen cyanide if heated or placed in contact with concentrated acids.

Clorox 2 powder may be irritating to mucous membranes. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Consult Material Safety Data Sheets for additional safety information.

Preparation

Prepare Instant Light crystals: Mix the dry ingredients thoroughly in a 400-mL beaker, but do not grind! Try mixing the luminol and potassium ferricyanide together first and then add the Clorox 2 with frequent stirring. Lastly, pour the dry mixture from one beaker to another beaker to ensure a homogenous mixture. Do this last step in an operating fume hood to reduce airborne dust.

Procedure

1. Fill a beaker with distilled or deionized water, and place the beaker on a magnetic stirrer. The size of the beaker and amount of water is up to you.
2. Turn off the lights in your classroom.
3. Add some Instant Light crystals to the water. About two teaspoons for every 200 mL water works well.
4. The blue chemiluminescent glow will begin instantly and last for several minutes.

Disposal

Allow the Instant Light solution to fully react (stir for 15 minutes) and then flush down the drain with excess water according to Flinn Suggested Disposal Method #26b.

Tips

- Clorox 2 is a commercial product. Due to formulation variations of the Clorox 2, actual times of the chemiluminescent glow may vary greatly from one batch to another. The glow should last from 90 seconds to 5 minutes depending on the formulation. To enhance the glow, make sure the room is completely dark.
- Adding a small amount of a fluorescent dye along with the luminol will produce different colors of light. Try small amounts (0.005 g) of disodium fluorescein (yellowish green) or Rhodamine B (red).
- Use hot or cold water to see how temperature affects the kinetics of the chemiluminescent reaction.
- Sprinkle Instant Light crystals on a wet towel; they will light up like stars.

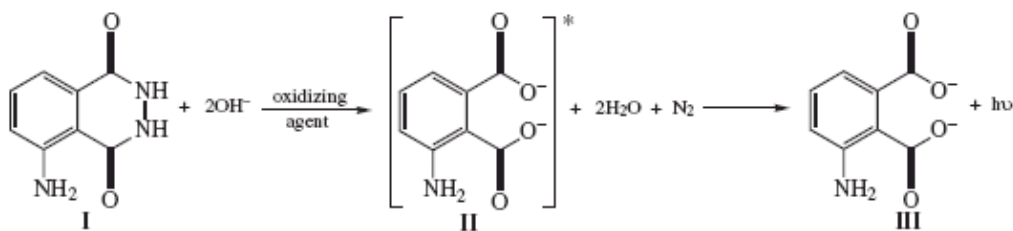
Discussion

Chemiluminescence is defined as the production or emission of light that accompanies a chemical reaction. Light emission results from the conversion of chemical energy into light energy due to changes in the composition of a chemiluminescent material.

The “flame test” colors observed when different metal salts are burned in a Bunsen burner flame are examples of a type of chemiluminescence known as pyroluminescence. The oxidation of luminol (3-aminophthalhydrazide) in this demonstration is another well-known example of chemiluminescence. The light-producing chemical reactions of luminol were discovered by H. O. Albrecht in 1928. Since that time numerous procedures have been developed to produce light using luminol. Experiments have shown that the following “ingredients” are necessary for luminol to exhibit chemiluminescence—a basic (alkaline) pH, an oxidizing agent, and a catalyst. In this demonstration, the oxidizing agent is Clorox 2, which also maintains the basic pH needed, and the catalyst is the iron(III) cation in potassium ferricyanide.

Oxidation of luminol and the resulting chemiluminescence occurs in the following sequence of reactions:

1. Chlorox 2 acts as a base and converts luminol (structure I) into a dianion.
2. Chlorox 2 oxidizes the dianion form of luminol to aminophthalate ion (structure II), which is produced in an excited electronic state (electrons not occupying their lowest energy orbital).
3. The excited aminophthalate ion decays to a lower energy ground state and gives off light in the process (structure III). The emitted light has a wavelength of 425 nm, which is in the blue region of the visible spectrum.



Electrons & Quantum Mechanics

Activity #3 - Flame Test Lab

Questions to be investigated

How do we identify metals?

Objectives

Students will observe and identify elements based on the color produced in a flame test.

Teacher Notes

Background information

- Remind students to wear safety goggles and follow all lab rules.
- To avoid having students walk around during the lab, put a set of solutions in small beakers or vials labeled at each lab table. There should be two lab groups per table and they should share the solutions. A student helper would distribute the unknowns, again to keep students from leaving their lab tables.
- In order for students to obtain maximum results, have the lights in the classroom dim. This is particularly helpful if the students are using a spectroscope or diffraction grating.

By heating the elements, the electrons in the atom become excited. When the electrons fall back to their ground state, they give off photons of light at characteristic energies. If the amount or abundance of an element present is altered, it is possible to change the intensity of the lines, (their brightness), because more photons would be produced. But it is not possible to change their characteristic colors - no matter how much or how little was present, the pattern of lines would be the same.

Gather materials. Platinum or nichrome wire may be used. Compounds with chloride as the anion may be substituted for the nitrates. For example, sodium chloride can be used instead of sodium nitrate. This lab can also be done using 6M HCl as the cleaning agent instead of the 1M. The student handout uses 1M for safety reasons. The solutions used can be 0.5 M or stronger.

Discussion and wrap up

Discuss the applications of flame tests. Ask students how else they might be useful. Discuss the limitations of flame tests. A test viewed only with the naked eye is limited by individual color interpretations and by ambiguities. For example, rubidium and cesium produce the same color as potassium. Blue cobalt glass can be used to filter out the light for a better color contrast.

Materials

Safety goggles

Test tube rack

Platinum or nichrome wire loop

Copula

50 ml beaker

Gas burner

8 test tubes

Optional cobalt-blue glasses

Chemicals:

6M HCl Hydrochloric acid (pour 150 mL concentrated HCl into 150 ml of distilled water)

10 g of each of the following:

Sodium Nitrate, NaNO_3

Lithium Nitrate LiNO_3

Potassium Nitrate, KNO_3

Copper II Nitrate, $\text{Cu}(\text{NO}_3)_2$

Calcium Nitrate, $\text{Ca}(\text{NO}_3)_2$

Barium Nitrate, $\text{Ba}(\text{NO}_3)_2$

Strontium Nitrate, $\text{Sr}(\text{NO}_3)_2$

Safety Concerns

Safety goggles

Hydrochloric acid is corrosive and can cause severe injury. If you spill acid on yourself, immediately flush the affected area with water for 2-3 minutes and notify the teacher. If acid should get in your eyes, begin flushing your eyes with water immediately and continue doing so for at least 20 minutes. If there is an eye wash fountain equipped with continuously running water in the laboratory, use it.

If acid is spilled on the laboratory bench or on the floor, neutralize the spill with solid sodium bicarbonate, NaHCO_3 , before wiping it up with sponges or paper towels. The acid has been neutralized when bubbles of gas no longer form after addition of the sodium bicarbonate.

Do not at any time touch the end of the wire loop used in the flame tests. The wire gets extremely hot and can cause severe burns. Remember that a wire can be hot and yet appear no different from a cool wire.

Real-World Applications

Fireworks

Street lights are sodium vapor lamps in many communities. These lamps have an orange yellow tint. You can see from the emission spectrum why the sodium vapor lamps would appear yellow and not white. These lamps consume less energy than the older blue colored mercury vapor lamps. Mercury vapor lamps have been sold in hardware stores for yard lighting.

One of the odd things that we sometimes notice is that colored things have a different appearance in natural daylight than they do under mercury vapor or sodium vapor lamps. This is reasonable because the daylight includes all of the wavelengths of white light and the vapor lights only emit a few specific colors that can be reflected into our eye off of any illuminated article.

Sources

Wilbraham, Antony C, and Dennis D. Saley. Chemistry Laboratory Manual Teacher edition. . Addison- Wesley, 1990.

Online lab:

<http://www.800mainstreet.com/spect/emission-flame-exp.html#Anchor-barium>

Background information and excellent pictures of flame test colors.

[Spectroscopy: Light & Element Identification](#)
[Flame Test](#)

Supporting NASA explores Article(s): ["Say Aaahhhh"](#)

Procedure

Identification of metals

As you perform the experiment, record your observations in the table following the procedure.

1. Obtain small samples of salts in labeled test tubes. For each sample, an amount of salt that just fills the tip of a scoopula is sufficient.
2. Pour approximately 15 mL of 6M HCl into a 50 mL beaker. Clean the wire loop by first dipping it into the 6M HCl and then heating it in the hot flame of a gas burner. Continue to dip and heat until no color comes from the wire as it is heated.
3. Dip the clean wire loop into a sample of metal salt and then heat the loop in the burner flame. Record the color of the flame in the data table. Test the remaining samples in the same way. Be certain to clean the wire loop thoroughly before each new sample is tested. Record your observations.
4. Observe the colors of the flames produced by heating NaNO_3 and KNO_3 . View the flames through cobalt glass. Record your observations.
5. Obtain an unknown salt from your teacher. Perform the flame test and record your observations.

Name _____ Date _____ Class _____

Identification of Metals: Flame Tests

Wilbraham, Antony C, and Dennis D. Saley. Chemistry Laboratory Manual Teacher edition. . Addison-Wesley, 1990.

Introduction

When elements are heated to high temperatures, some of their electrons are excited to higher energy levels. These excited electrons can then fall back to lower energy levels, releasing the excess energy in packages of light called photons, or light quanta. The color of the emitted light depends on its energy. Blue light is more energetic than red light, for example. When heated, each element emits a characteristic pattern of light energies which is useful for identifying the element. The characteristic colors of light produced when substances are heated in the flame of a gas burner are the basis of flame tests for several elements.

In this experiment, you will perform flame tests for several metallic elements.

Objectives

1. To observe the colors emitted by various metal ions.
2. To evaluate flame testing as a method of detection of metals.

Equipment

Safety goggles
Platinum or nichrome wire loop
50 ml beaker
8 test tubes

Test tube rack
Copula
Gas burner
Optional cobalt-blue glasses

Materials

6M HCl Hydrochloric acid
Sodium Nitrate, NaNO_3
Potassium Nitrate, KNO_3
Calcium Nitrate, $\text{Ca}(\text{NO}_3)_2$

Strontium Nitrate, $\text{Sr}(\text{NO}_3)_2$
Lithium Nitrate LiNO_3
Copper II Nitrate, $\text{Cu}(\text{NO}_3)_2$
Barium Nitrate, $\text{Ba}(\text{NO}_3)_2$

Safety

1. Wear safety goggles
2. Hydrochloric acid is corrosive and can cause severe injury. If you spill acid on yourself, immediately flush the affected area with water for 2-3 minutes and notify the teacher. If acid should get in your eyes, begin flushing your eyes with water immediately and continue doing so for at least 20 minutes. If there is an eye wash fountain equipped with continuously running water in the laboratory, use it.

If acid is spilled on the laboratory bench or on the floor, neutralize the spill with solid sodium bicarbonate, NaHCO_3 , before wiping it up with sponges or paper towels. The acid has been neutralized when bubbles of gas no longer form after addition of the sodium bicarbonate.

3. Do not at any time touch the end of the wire loop used in the flame tests. The wire gets extremely hot and can cause severe burns. Remember that a wire can be hot and yet appear no different from a cool wire.

Procedure

As you perform the experiment, record your observations in the table following the procedure.

1. Obtain small samples of salts in labeled test tubes. For each sample, an amount of salt that just fills the tip of a scoopula is sufficient.
2. Pour approximately 15 mL of 6M HCl into a 50 mL beaker. Clean the wire loop by first dipping it into the 6M HCl and then heating it in the hot flame of a gas burner. Continue to dip and heat until no color comes from the wire as it is heated.
3. Dip the clean wire loop into a sample of metal salt and then heat the loop in the burner flame. Record the color of the flame in the data table. Test the remaining samples in the same way. Be certain to clean the wire loop thoroughly before each new sample is tested. Record your observations.
4. Observe the colors of the flames produced by heating NaNO_3 and KNO_3 . View the flames through cobalt glass. Record your observations.
5. Obtain an unknown salt from your teacher. Perform the flame test and record your observations.

Flame tests observations table

Ion	Flame color/observations
Sodium, Na ⁺	
Potassium, K ⁺	
Calcium, Ca ⁺	
Barium, Ba ²⁺	
Strontium, Sr ²⁺	
Lithium, Li ⁺	
Copper, Cu ²⁺	
Optional: Sodium, Na ⁺ Cobalt glass	
Optional: Potassium, K ⁺ Cobalt glass	
Unknown	

Results and Conclusions

1. List the elements that gave the most easily identified colors. Which elements are least easily identified?

2. Which element gives the most intense color?

3. What does this tell you about the amplitude

4. Do you think that flame tests are valuable for detecting metal ions present in a mixture of metal ions? Explain.

5. The energy of colored light increases in the order red, yellow, green blue, violet. List the metallic elements used in the flame tests in increasing order of the energy of light emitted.

6. What is the purpose of using the cobalt glass in the identification of sodium and potassium?

7. Consider the colors of the flames produced by various metals ions. Do you see any relationship between the energies of light emitted and the positions of elements in the periodic table?

Going further

Paper logs soaked in solutions of metal salts and dried are sold for producing colored flames in fireplaces. Make a list of ingredients on logs at a sales outlet to see if you have predicted correctly.

Name _____ Date _____ Class _____

Experimental Procedure

Student Sheet

Part 1 Flame tests and identification of an unknown metal.

Observe and record the color of the flame for each metal ion. Remember the metal ions are paired with a nonmetal ion in an ionic formula unit. The electrical charges have to add to zero. The metal ions are converted to atoms in the flame and then excited by the heat from the Bunsen burner flame. The nonmetal ions, anions, do not get converted to atoms and do not and emit visible light like the metals do.

Repeat procedure for each known. Record the color observed for the unknown and use the color match to identify the metal atom that is the produced from the cation in the unknown.

Metal ion	return to procedure	Observed Flame color
barium	click for flame test	_____
calcium	click for flame test	_____
sodium	click for flame test	_____
rubidium	click for flame test	_____
potassium	click for flame test	_____
lithium	click for flame test	_____

Part 1 Flame tests for unknown elements

Unknowns		Flame color	Identity of metal ion based on flame test
Unknown 1	click for flame test	_____	_____
Unknown 2	click for flame test	_____	_____

Part 2 Observing line spectra with the spectroscope

In the second part of the experiment you will observe the color of light emitted by excited gases of elements in sealed glass tubes called "spectrum" tubes. Direct current, DC, high voltage electrons are used to excite the atoms in the spectrum tube. High voltage means 1000 to 2000 volts. This is more than 10 times normal household voltage which is 120 volts AC.

The excited atoms release the energy they gained. Some of this energy is in the form of heat and some is in the form of light. The billions of excited atoms release energy. Each excited atom releases a single pulse of light energy as it returns to the "ground" state or low energy state. There are so many pulses emitted the light appears to be continuous.

The excited atoms do not all emit the same energy light because the amount of energy that excited them may differ, but there are limitations on the colors they do emit. The kind of light depends on the size of the gaps between the "shells" or energy levels in the atom. The electrons are changing "n" values in the atom. Remember "n" can have only positive whole number values like 1, 2, 3 ... up to infinity.

The kind of light energy that can be emitted by excited atoms is unique for an element. The pattern of "lines" or colors emitted can be used to identify an element. A powerful extension of this is the ability to measure amounts of an element by measuring the brightness of the emitted light.

A spectroscope can separate the light produced by an emission tube. The color seen by the naked eye is a combination of a number of colors of light. These are separated by a prism or a diffraction grating which acts like a prism. The emission lines can be seen when you look through the spectroscope at the light source. You will be able to observe the "line" spectrum for the elements and record the spectral lines.

Part 2 Emission line spectra for selected elements

Element	Emission	Emission spectrum
Sodium	click here to view emission tube	click here to view emission spectrum
Neon	click here to view emission tube	click here to view emission spectrum
Mercury	click here to view emission tube	click here to view emission spectrum
Helium	click here to view emission tube	click here to view emission spectrum

Questions and observations

Examine the spectra for the elements Na, Ne, Hg or He and answer the following questions. Fill in the following table with your answers.

How do these emission spectra compare in terms of colors and numbers of emission line positions?

Are the spectra identical?

What if anything is similar?

What is different?

FILL IN THE FOLLOWING TABLE WITH YOUR ANSWERS

Element with greatest number of visible emission lines	_____
Longest wavelength in the spectrum of this atom in nanometers	_____
Color of light for this longest wavelength	_____

What suggestions do you have additions or changes to this experiment?

What "new" idea did you learn from this experiment?

Why does a sodium vapor street light look yellow instead of white?

What would you expect to happen to the size (volume) of a hydrogen atom when the outer electron moves from the $n = 2$ shell to the shell with $n = 4$? Would the volume increase?

Electrons & Quantum Mechanics

Activity #4 – Investigation s becomes before d

Questions to be investigated

How is it possible for a transition metal to have more than one oxidations state?

If you have the two of the same transition elements with different oxidation states, which oxidation state is more stable?

Objectives

Students will observe two forms of ions.

Students will describe the relationship between electron configuration and oxidation number for these two ions.

Teacher Notes

DISPOSAL: Collect and recycle the FeCl_3 by reusing it in the laboratory. Mix the HCl with water and a large amount of Na_2CO_3 and $\text{Ca}(\text{OH})_2$. When neutralized flush down the drain with a large amount of water.

Materials

Each group of students need:

2 large test tubes

test tube rack

2 g iron filings or steel wool

2 g iron (III) chloride, FeCl_3

4 ml concentrated HCl

12 ml distilled water

2 rubber stoppers (fit into test tubes)

balance

graduated cylinder

goggles (each student)

aprons

Safety Concerns

Goggles are needed for every student; aprons if available.

CAUTION: Always add acid to water. Concentrated HCl causes severe burns; avoid skin contact. Rinse spills with plenty of water. Dispose of all chemicals as instructed above.

CAUTION: FeCl_3 is a skin irritant.

Sources

Smoot, Robert C., and Richard G. Smith. Chemistry A Modern Course Teacher Resource Book. Columbus OH: Merrill, 1987.

Procedure

1. Into the first large test tube, place approximately 2 g of iron filings. Add 4 ml of water to the test tube, followed by 4 ml of concentrated HCl. Record your observations.
2. Into a second large test tube, place approximately 2 g of iron (III) chloride, FeCl_3 . Add 8 ml of water to the test tube. Place a rubber stopper into the test tube and shake until some of the compound has dissolved. Record your observations.

Name _____ Date _____ Class _____

Investigation s becomes before d Student Sheet

Questions to be investigated

How is it possible for a transition metal to have more than one oxidations state?

If you have the two of the same transition elements with different oxidation states, which oxidation state is more stable?

Objectives

Students will observe two forms of ions.

Students will describe the relationship between electron configuration and oxidation number for these two ions.

Materials

2 large test tubes	2 rubber stoppers (fit into test tubes)
test tube rack	balance
2 g iron filings or steel wool	graduated cylinder
2 g iron (III) chloride, FeCl_3	goggles (each student)
4 ml concentrated HCl	aprons
12 ml distilled water	

Safety

Goggles must be worn during this experiment; aprons if available.

CAUTION: Always add acid to water. Concentrated HCl causes severe burns; avoid skin contact. Rinse spills with plenty of water. Dispose of all chemicals as instructed above.

CAUTION: FeCl_3 is a skin irritant.

Procedure

1. Into the first large test tube, place approximately 2 g of iron filings. Add 4 ml of water to the test tube, followed by 4 ml of concentrated HCl. Record your observations.
2. Into a second large test tube, place approximately 2 g of iron (III) chloride, FeCl_3 . Add 8 ml of water to the test tube. Place a rubber stopper into the test tube and shake until some of the compound has dissolved. Record your observations.

Analysis and Conclusions

1. Record your observations below.

a. Test tube 1 _____

b. Test tube 2 _____

2. Write the electron configuration for the iron metal.

3. Which two electrons are lost to form the pale green Fe^{2+} ion?

4. Which additional electron is lost to form the yellow Fe^{3+} ion?

5. Other transition elements exhibit oxidation states from 2+ up to 8+. Why does iron have only the 2+ and 3+ oxidation states?

6. Would you predict iron to be more stable at the 2+ or 3+ oxidation states? Explain your answer.

Analysis and Conclusions **answers**

- Record your observations below.
 - Test tube 1 ___pale green solution forms _____
(the Fe^{2+} forms when the acid oxidizes the metal)
 - Test tube 2 ___yellow solution forms _____
(Fe^{3+} dissociates in water)
- Write the electron configuration for the iron metal.
_____ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ _____
- Which two electrons are lost to form the pale green Fe^{2+} ion?
_____ $4s^2$ _____
- Which additional electron is lost to form the yellow Fe^{3+} ion?
_____ the paired 3d electron _____
- Other transition elements exhibit oxidation states from 2+ up to 8+. Why does iron have only the 2+ and 3+ oxidation states?
_____ The empty 4s and half-filled 3d sublevel form stable ions. _____
- Would you predict iron to be more stable at the 2+ or 3+ oxidation states? Explain your answer.
_____ The 3+ state is more stable because it results in a half filled sublevel. _____

Electrons & Quantum Mechanics

Activity #5 – Demonstrations

Questions to be investigated

What do s and p orbitals look like?

How are wavelength and color related?

Objectives

Demonstrate a variety to show wave formation, s and p orbitals and relate real world examples.

Sources

http://dwb4.unl.edu/chem_source_pdf/ChemSource.htm

Procedure

See document

Concept/Skills Development



http://dwb4.unl.edu/chem_source_pdf/ChemSource.html

Demonstrations

Demonstration 1: Spectra in the Real World

Give each student a small square of plastic diffraction grating to view H_2 , He, Ne, etc. emission from spectrum tubes. Let students take the grating home to view sodium street lamps and mercury vapor lamps. WARNING: Do not look at the sun!

Demonstration 2: Getting a Charge out of Things

If students haven't seen the demonstration of an amber or glass rod (or a length of PVC pipe works very well) rubbed with fur, silk, or hair, and then used to attract or repel pith balls or other objects (such as suspended balloons), this is a good place to perform it.

Demonstration 3: Standing Waves

An old-style man's electric razor, with the cutting top removed, has an exposed vibrating post. By attaching a string to this post, you can show quantized standing waves when the razor is turned on.

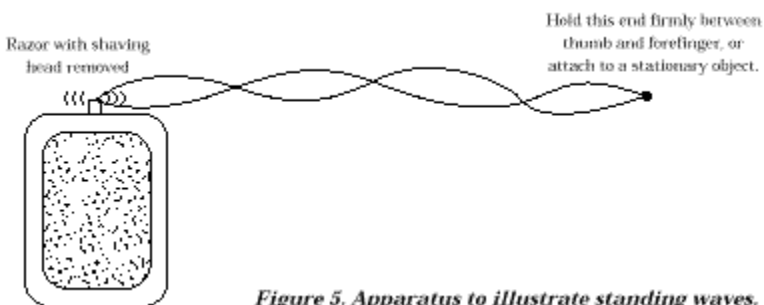


Figure 5. Apparatus to illustrate standing waves.

Demonstration 4: Modeling Orbitals.

Spherical balloons can be used to model s-orbitals. Two spherical or oblong balloons tied together can be used to model p-orbitals.

Materials

6 Balloons (2 each of 3 different colors). Tie pairs of balloons (2 the same color) together. Bunch.

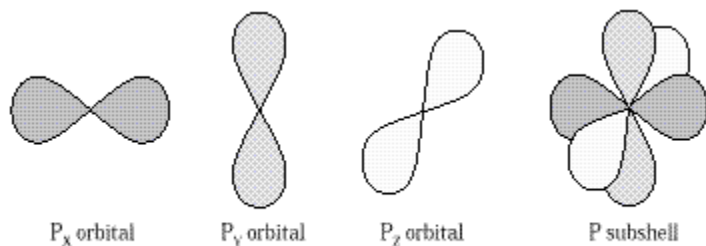


Figure 6. Balloons illustrating p-orbitals.

Demonstration 5: Wavelength and Color

To demonstrate the relationship between wavelength and color in the visible range from red to violet, use a spectrophotometer. Place a piece of white chalk in a spectrophotometer cell (see Figure 7). Cut (or sandpaper) the chalk at an angle that will reflect light upward to the student's eye looking down into the cell holder. Open the slit width. Rotate the wavelength knob to compare colors and wavelengths.

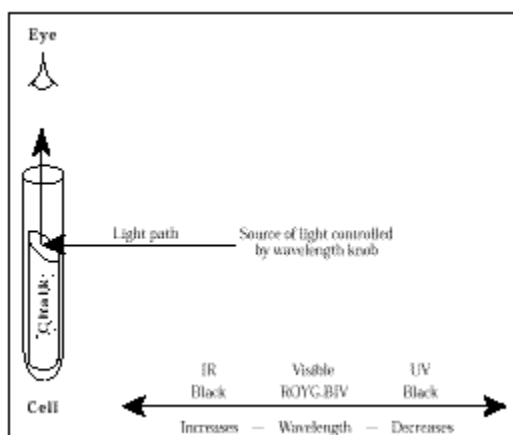


Figure 7. Demonstrating the relationship between color and wavelength.

Demonstration 6: Simple Mass Spectroscopy

Materials

Overhead projector

Sheet of Plexiglas, 8-1/2 x 11 in

Magnet, disk shape, 3-4 in diameter

3-4 Steel ball bearings of various sizes

Procedure

1. Form a shallow inclined plane with the Plexiglas sheet on top of the overhead projector, by using a small block of wood or other material to elevate one end of the sheet.
2. Insert the magnet underneath the center of the inclined plane.
3. Roll the steel balls down the incline. If they are rolled near but not over the magnet (they will stop) their direction of motion will be altered. The smaller ball bearings will be deflected through a greater angle than the larger ones.

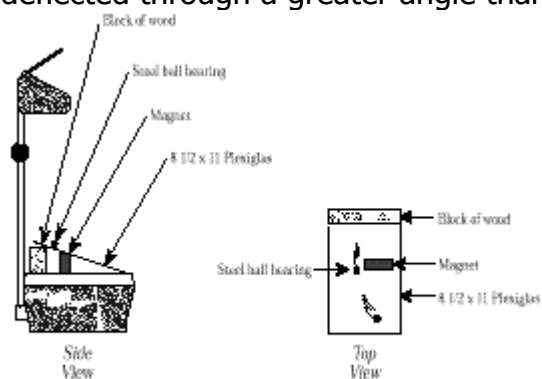


Figure 8. Simple mass spectroscopy demonstration.

Demonstration 7: Flame Tests

Absorption and emission spectra are used to identify any elements. When elements are heated to high temperatures they may be placed in an excited state; in this excited state valence electrons move to higher energy levels. When the electrons return to their ground state, they may emit visible light of characteristic colors that can be used to identify the element. In this demonstration, students will identify the colors of the emission spectra of some metallic ions, using a burner and wood splints. Place the substance in the hottest part of the flame

Materials

Burner

Wood splints

Copper(II) nitrate, $\text{Cu}(\text{NO}_3)_2$, solid crystals

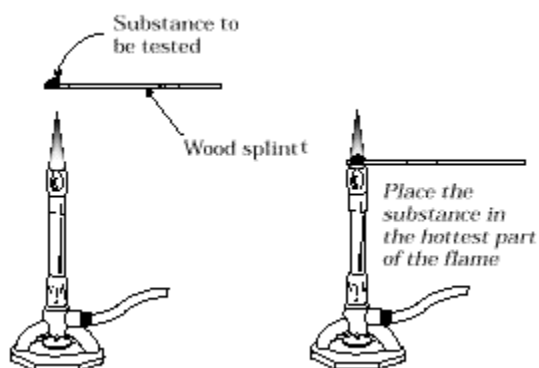
Barium nitrate $\text{Ba}(\text{NO}_3)_2$, solid crystals

Sodium nitrate, NaNO_3 , solid crystals

Strontium nitrate, $\text{Sr}(\text{NO}_3)_2$, solid crystals

Procedure

Place a small amount of one of the solid ionic compounds (about the size of a rice grain) on the tip of a wooden splint. Place the splint at the edge of the hottest part of the flame (top of inner blue cone) and observe the color. (Some students, particularly some males, may be color blind. Consider allowing students to compare their observations in small groups, alerting them to the possibility of color blindness.) (Note: Good results have been obtained with concentrated syrups of the nitrate salts.)



Results

Key Questions

1. What are the principal parts of an atom?
2. What forces hold atoms together? []
3. How are electron orbitals different from orbits?
4. What is an energy level diagram?
5. How are electron energies related to orbitals?
6. How do cations form?

7. What determines the ionization energy of an element?
8. What determines the commonly formed cation of an atom?
9. What is the process of electron excitation?
10. Why do atoms emit light when excited?
11. How do the physical sizes of atoms and nuclei compare?
12. What are the three dimensional shapes of s and p orbitals?
13. What is meant by wave-particle duality?

14. What information is provided by quantum numbers?
15. How do quantized and continuous processes differ?
16. What is the electron configuration of (specify an element)?

Answers to Key Questions

1. The nucleus and electron cloud. The nucleus contains protons, neutrons and many other subnuclear particles (see Nuclear Chemistry module). The electrons can be divided into core and valence electrons.
2. Coulombic (electrostatic) attractions between oppositely charged protons and electrons.
3. Orbitals refer to the probable location or region of space (a volume) where an electron is likely to be found with a specified degree of certainty. It arises from application of wave and uncertainty considerations. Orbits are Niels Bohr's representation of rigid circular or elliptical paths followed by particulate electrons.
4. A graphic representation of the relative energies possessed by electrons in various orbitals as described by their quantum numbers.
5. The principal quantum number gives a general indication of an electron's energy and sets restrictions on its orbital type. Higher energy orbitals are those with higher probability of being far from the nucleus.
6. If an electron absorbs energy in excess of the amount corresponding to the highest energy orbital in its electron cloud, it will be ionized, that is, leave the atom. The resulting positive ion left behind is a cation.
7. The nuclear charge (the binding force's source) and the energy possessed by the electron (its energy level) combine (actually compete) to determine additional (ionization) energy required.
8. When a metal and nonmetal combine, the overall process is exothermic, largely due to the highly negative lattice energy involved; thus reaction takes place. One of the processes that requires energy in forming the compound is the ionization of the metal ion. It is the relative magnitudes of the ionization energies of a particular metallic species that determines the charge on the cation. For example, sodium has a first ionization energy (I_1) of 495.8 kJ/mol; its second ionization energy (I_2 , the energy needed to remove a second electron from sodium) is 4562.4 kJ/mol, almost 10 times I_1 ! No reaction involving the formation of Na_2^+ could ever be expected to be exothermic; hence, Na_2^+ does not exist. A second example is the formation of magnesium ions. I_1 , I_2 , and I_3 for magnesium are 737.7, 1450.7, and 7732.6 kJ/mol, respectively. The large energy requirement for the formation of Mg_3^+ precludes its existence. Mg^+ requires the least amount of energy for formation, but is unstable because of its half-filled 3s orbital. Hence, the most stable magnesium ion is Mg^{2+} , requiring a total energy expenditure of 2188.4 kJ/mol, which is certainly within the realm of a possible overall exothermic process for the formation of a magnesium salt.
9. An electron is said to be excited if it absorbs energy and moves to a higher energy orbital.
10. Being in an excited state means having excess energy. Like anything else, atoms tend to lose their excess energy, so the electron drops to a lower energy orbital. The lost energy is given off as light ($DE = h\nu$).
11. Atoms tend to have diameters of around 0.1 nm (1-10 Å). The nucleus is about 10^{-15} m (1 femtometer) in diameter but contains 99.9% of the mass.]

12. The 90% probability boundary for an s-orbital is a sphere. For p-orbitals the boundary is dumbbell shaped. Three p-orbitals, one dumbbell per axis, comprise a set.
13. In some instances, the behavior of electromagnetic radiation and electrons can best be understood in terms of wave theory, and in other instances, their behavior can best be understood if we treat them as discrete particles. Electromagnetic radiation is usually described in terms of wavelength or frequency; the quantity most closely associated with particles is mass or momentum, the mass-velocity product. Louis de Broglie in 1924 offered the startling proposition that light may sometimes display particle-like characteristics, and that small particles may sometimes display wavelike properties. He summarized his proposition in a now famous equation $\lambda = h/mv$, where λ is the wavelength, h is the Planck constant, m is the mass, and v is the velocity of the object in question. Wave-particle duality is only important when the wavelength is of atomic dimensions, i.e., of the order of $10^1 - 10^3$ picometers.
14. All we can hope to know about an atomic electron: its energy, angular momentum, magnetic moment and spin.
15. Anything quantized can be described in terms of separate discrete units. Those that are continuous are without beginning or end and may be had in any quantity. Ice cubes are quantized but water flows continuously.
16. Examples: H $1s^1$; Li $1s^2 2s^1$; Ga $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$

Counterintuitive Examples

1. Adjustable light dimmer switches (unless specifically designed) do not work on fluorescent lamps. You can demonstrate this fact by first connecting a standard, preferably clear, incandescent light bulb to a dimmer switch or laboratory voltage controller such as a Powerstat. By slowly increasing the voltage, you can brighten the incandescent bulb in a continuous manner. Repeat using a fluorescent lamp. Since the mercury gas enclosed in the fluorescent lamp's envelope requires a minimum ionization energy, a quantized process, the lamp will not go on until this minimum energy is applied.
2. A weight attached to a string and rotated around one's head continues in motion only as long as energy is applied. An electron, perceived to be in a classic Bohr orbit, does not slow down. This behavioral discrepancy was one of the counterintuitive problems that Bohr had to solve.
3. Energy changes do not always seem to be quantized. If you heat a piece of metal, its temperature rises on a continuum; the metal does not seem to acquire energy in quantum leaps. On the other hand, if you step halfway between two ladder rungs, you find that you remain on the first step; you must acquire potential energy in quanta dictated by the step-size of the ladder. The manner of acquiring energy in these two instances is not different. The energy step-sizes for a metal are so infinitesimally small that they cannot be detected by a relatively crude probe such as a thermometer.
4. As electrons are added to atoms in the same period of the Periodic Table, the atomic size decreases. Adding an electron also entails adding a proton to the nucleus, thus increasing effective nuclear charge. The nuclear charge density increase is proportionally greater than the electron energy increase, leading to more compact (smaller) atoms. A Periodic Table showing atomic sizes (such as the one included in the Periodicity module) will help illustrate this point.
5. Light radiated by an excited atom consists of several colors, but the human eye perceives only one composite color. This composite color can be resolved into its components by viewing it through a diffraction grating. A spectral tube for neon or helium, or even an orange night light, illustrates this fact very nicely.
6. Because electron probability functions are asymptotic, they do not become zero even at a great distance from the atom's nucleus. Consequently, there is a very small chance that some electron associated with an atom in your body is at this moment very far away. Perhaps a part of you has been to the moon and you just don't know it!

Electrons & Quantum Mechanics

Activity #6- Complex Ions and the Spectrochemical Series

Questions to be investigated

What is the spectrochemical series of complex ions?

Objectives

Students will develop the spectrochemical series by observing the colors of several complex ions.

Teacher Notes

This experiment is appropriate for an Advanced Placement course

Time required

One to three lab periods depending on whether the solutions are made up by the teacher or the students.

Modifications/Substitutions

1. Copper sulfate pentahydrate is available from a garden supply store as root killer.
2. Possible substitutions for the ligands are: hydrochloric acid--muriatic acid available from a hardware store, ammonia solution--household ammonia available from a grocery store, sodium thiosulfate-- photo fixer from a camera/photo supply store, oxalic acid--ZUD from a grocery store, acetic acid--vinegar from a grocery store, and tartaric acid--cream of tartar from a grocery store.

Materials

Chemicals

copper sulfate pentahydrate	oxalic acid
nickel sulfate heptahydrate	acetic acid
cobalt sulfate heptahydrate	tartaric acid
95% ethyl alcohol	EDTA
possible sources of ligands:	sodium bromide
hydrochloric acid	ethylenediamine
ammonia	dimethylglyoxime
sodium thiosulfate	

Equipment

pH paper
small test tubes
eyedroppers

Safety Concerns

Avoid skin contact with all the solutions; concentrations used in this experiment may cause irritation. Goggles must be worn throughout the experiment.

Sources

<http://web.archive.org/web/20050210092308/www.woodrow.org/teachers/chemistry/institutes/1986/exp35.html>

Procedure

1. Prepare the following cation test solutions: 1.0 M CuSO_4 (dissolve 250 g $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in sufficient distilled or deionized water to make 1.0 L of solution), 0.10 M NiSO_4 (dissolve 28.1 g of $\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$ in sufficient distilled or deionized water to make 1.0 L of solution), and 0.10 M CoSO_4 (dissolve 28.1 g $\text{CoSO}_4 \cdot 7\text{H}_2\text{O}$ in sufficient water to make 1.0 L of solution).
2. Prepare saturated solutions of tartaric acid and of oxalic acid.
3. Adjust the pH of the solutions prepared in step 2 so that they are neutral to pH paper by adding small amounts of ammonia solution.
4. Prepare a 1% solution of dimethylglyoxime in ethyl alcohol solution.
5. Measure 1 ml of the copper(II) solution into 11 small test tubes. Repeat with the cobalt(II) and nickel(II) solutions.
6. Set aside a test tube of each metallic cation as a control.
7. Add several drops of each ligand solution or a small scoop of a solid ligand to each of the remaining 10 test tubes containing the copper(II) ion. Mix the contents of each tube thoroughly after each addition. Repeat for the cobalt and nickel solutions. Observe carefully as the ligands are added. Continue to add ligand until a noticeable color change occurs. If no color change has been observed after an excess of ligand (approx. 3 mL) has been added, record no visible change on the data chart. Always compare the sample with the control. Note: Some of the complexes may have several intermediate states in which the color may vary; therefore, be sure to add an excess of ligand after the initial color change.
8. Line up the resulting complexes for each cation in order of the spectrum (ROYGBIV).
9. Record observed color changes in a data table similar to the one following.

	Copper+2		Cobalt+2		Nickel+2	
Ligand Added	Initial Color	New Color	Initial Color	New Color	Initial Color	New Color

10. Determine the order of the spectrochemical series using information in the discussion below.

Assessment Ideas

Have students write up a formal lab write up before beginning the lab. In addition, have students' research ligands.

Name _____ Date _____ Class _____

Complex Ions and the Spectrochemical Series

Purpose

The purpose of this experiment is to develop the spectrochemical series by observing the colors of several complex ions.

Description

Students will develop the spectrochemical series by investigating the colors of the complex ions formed when solutions of various ligands are added to aqueous solutions of copper(II), nickel(II), and cobalt(II) ions.

Materials

Chemicals

copper sulfate pentahydrate	oxalic acid
nickel sulfate heptahydrate	acetic acid
cobalt sulfate heptahydrate	tartaric acid
95% ethyl alcohol	EDTA
possible sources of ligands:	sodium bromide
hydrochloric acid	ethylenediamine
ammonia	dimethylglyoxime
sodium thiosulfate	

Equipment

pH paper
small test tubes
eyedroppers

Hazards

Avoid skin contact with all the solutions; concentrations used in this experiment may cause irritation. Goggles must be worn throughout the experiment.

Procedure

1. Prepare the following cation test solutions: 1.0 M CuSO_4 (dissolve 250 g $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in sufficient distilled or deionized water to make 1.0 L of solution), 0.10 M NiSO_4 (dissolve 28.1 g of $\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$ in sufficient distilled or deionized water to make 1.0 L of solution), and 0.10 M CoSO_4 (dissolve 28.1 g $\text{CoSO}_4 \cdot 7\text{H}_2\text{O}$ in sufficient water to make 1.0 L of solution).
2. Prepare saturated solutions of tartaric acid and of oxalic acid.
3. Adjust the pH of the solutions prepared in step 2 so that they are neutral to pH paper by adding small amounts of ammonia solution.
4. Prepare a 1% solution of dimethylglyoxime in ethyl alcohol solution.
5. Measure 1 ml of the copper(II) solution into 11 small test tubes. Repeat with the cobalt(II) and nickel(II) solutions.
6. Set aside a test tube of each metallic cation as a control.
7. Add several drops of each ligand solution or a small scoop of a solid ligand to each of the remaining 10 test tubes containing the copper(II) ion. Mix the contents of each tube thoroughly after each addition. Repeat for the cobalt and nickel solutions. Observe carefully as the ligands are added. Continue to add ligand until a noticeable color change occurs. If no color change has been observed after an excess of ligand (approx. 3 mL) has been added, record no visible change on the data chart. Always compare the sample with the control. Note: Some of the complexes may have several intermediate states in which the color may vary; therefore, be sure to add an excess of ligand after the initial color change.
8. Line up the resulting complexes for each cation in order of the spectrum (ROYGBIV).
9. Record observed color changes in a data table similar to the one following.

	Copper + 2		Cobalt + 2		Nickel + 2	
Ligand Added	Initial Color	New Color	Initial Color	New Color	Initial Color	New Color

10. Determine the order of the spectrochemical series using information in the discussion below.

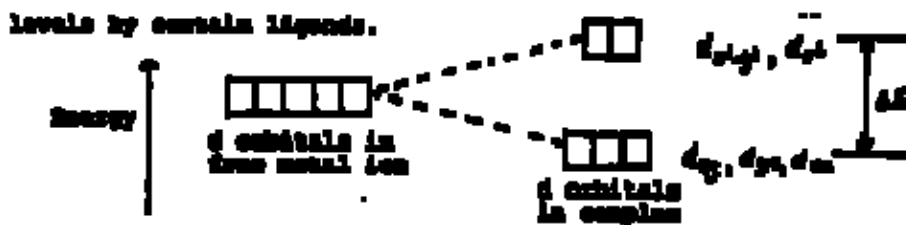
Disposal

Solutions may be flushed down the drain with ample water.

Discussion

In water solutions, transition metal ions are usually not found as single ions but as complex ions in which the metal ion is bonded to 2, 4, or 6 water molecules with coordinate covalent bonds. If the water solution is mixed with another species which can donate electrons to the metal, a new complex ion with the new species may be formed. The various species which can complex with metal ions are referred to as ligands. Often the color of the new complex ion is different from the original color of the metal ion in water.

The color of the complex arises from the crystal field splitting of the five degenerate d orbitals into 2 energy levels by certain ligands.



The energy gap, labeled ΔE , between the d orbitals is on the same order of magnitude as the energy of a photon of visible light. The actual magnitude of ΔE for a given complex and therefore the color of the complex is determined by the following: which ligand is used; which metal is used and the oxidation state of the metal. Thus, the color of the complex ion can be used to determine the wavelength of light absorbed by using the color rosette shown.



Source: Shakhshiri

The colors are arranged in a circle so that absorption of one color allows the color which is opposite on the rosette to become the visible color observed. For example, if an object appears to have a yellow color it is absorbing deep blue light; a red object is absorbing green light; etc. The wavelengths of light corresponding to their colors are also shown on the rosette. Of the wavelengths shown, 400 nm corresponds to the greatest energy and 720 nm corresponds to the smallest energy.

By determining the color absorbed by a series of ligands and the corresponding energy involved, the ligands can be ranked according to strength based on the relative amounts of energy absorbed (ΔE). Such a ranking is called a spectrochemical series. The following version of the series shows the ligands in order of increasing ΔE .

iodide ion < bromide ion < chloride ion < thiocyanate ion < fluoride ion < hydroxide ion < acetate ion < oxalate ion < water < ammonia < ethylenediamine < sulfite ion < nitrite ion < cyanide ion

The results obtained in this experiment for the ligands used with the copper(II), nickel(II), and cobalt(II) ions generally agree with the series as given above. Consumer chemicals, when available, were used as ligands. The weak acid solutions must be neutralized or made slightly basic with ammonium hydroxide before use as a ligand.

Possible questions and/or extensions include having students write formulas for the complexes formed, give names of the complexes formed, and correlate the strengths of the ligands as shown by the spectrochemical series results with the instability constants of the complex ions. Note, the strength of the ligand is not the only determining factor for instability constants; however, bidentate and tridentate ligands are generally bonded to the metal more tightly than monodentate ligands.

References

Beren and Brady, *Laboratory Manual for General Chemistry Principles and Structure*, Wiley and Sons, New York, 1986. A similar experiment is described.

Brady and Humiston, *General Chemistry, Principles and Structure*, Wiley and Sons, New York, 1986. Crystal field theory and complex ions are discussed.

Brown and LeMay, *Chemistry: The Central Science*, Prentice-Hall, Englewood Cliffs, NJ, 1981, p. 723. Crystal field theory and complex ions are discussed.

Shakhashiri, B.Z., *Chemical Demonstrations, Volume 1*, University of Wisconsin Press, Madison, 1983, p. 260. This work describes the theory of the color of complexes.

Name: _____ Date: _____ Period: _____

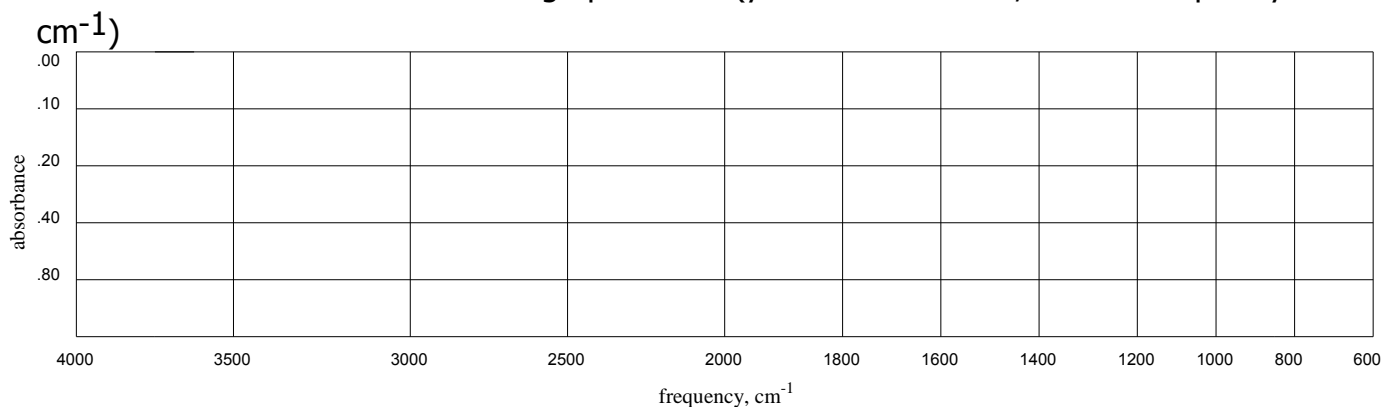
INFRARED SPECTROPHOTOMETRY ACTIVITY - VERSION B

Judith Ann Flint Baumwirt, Chemistry Instructor, Granada Hills High School/CSUN Math, Science and Technology Magnet

You are a forensic chemist and have been given a sample of a compound found at a crime scene. You have been asked to identify the substance. Once the sample has been prepared, properly transferred to a KCl cuvet and run through the IR Spectrophotometer, the following plot has been identified:

<u>Absorbance</u>	<u>Frequency, cm⁻¹</u>	<u>Absorbance</u>	<u>Frequency, cm⁻¹</u>
0.10	4000 (baseline)	0.15	1450
0.10	3000	0.40	1390
0.80	2950	0.10	1350
0.40	2900	0.10	1150
0.79	2850	0.90	1125
0.10	2800	0.15	1075
0.10	1500	0.10	1000
0.40	1480	0.10	800 (baseline)

Plot the above data on the graph below (y axis=absorbance, x axis=frequency

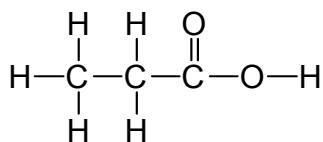


Utilizing the table below of known IR Spectra for specific molecular structures, identify the peaks in your plot above.

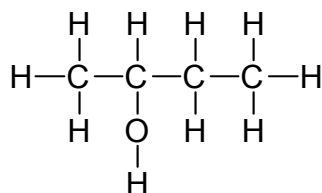
ν , cm^{-1}	Intensity *	Structure
1050-1400	(s)	C—O
1315-1475	(m-s)	C—H
1690-1750	(s)	C=O
1700-1725	(s)	HO—C=O
1770-1820	(s)	Cl—C=O
2100-2260	(m)	C≡C
2500-3000	(s)(vb)	O—H (in O=C—OH)
2800-3000	(m-s)	C—H
3200-3650	(s)	O—H

*Intensities: (s) = strong, (m) = medium, (w) = weak, (b) = broad, (vb) = very broad.

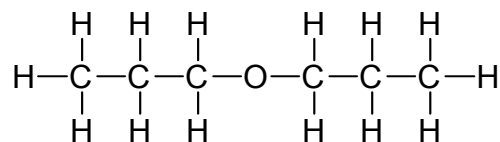
Based on this information determine the identity of your unknown substance based on the following choices: (Circle your answer)



propionic acid



sec-butyl alcohol



di-*n*-propyl ether

Name _____ Period _____ Date _____

Patterns in Electron Configuration

One of the many patterns contained in the periodic table is that of electron configuration. In this activity, you will identify these patterns. Later, you will use these patterns to determine the order in which electrons fill the orbitals of an atom. As you complete the activity, keep the following in mind:

- Period = row, Group = column
- Use the table on your book cover, which shows only valence electrons.
- There are two number systems for the Groups. We will focus on the A/B system.

1. Which Groups have an s-orbital as the *last* orbital? _____
2. Which Groups have a p-orbital as the *last* orbital? _____
3. Which Groups have a d-orbital as the *last* orbital? _____
4. Which section of the table is left? This section corresponds to the f-orbitals.
5. Look at Group 1A. What is the relationship between the Period number and the energy level of the valence electrons?
6. Look at Group 3A. What is the relationship between the Period number and the energy level of the valence electrons?
7. Look at Group 3B. What is the relationship between the Period number and the energy level of the d-orbitals?
8. Look at the Inner Transition Metals (bottom section). The Lanthanide series (58-71) is part of Period 6. The Actinide series (90-103) is part of Period 7. What is the relationship between the Period number and the energy level of the f-orbitals?
9. Look at all of the A Groups. What is the relationship between the Group number (1A, 2A, etc.) and the total number of valence electrons for each element? (Add up the exponents to find the total number of valence electrons.)