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Study and Testing Tips

What to study in order to prepare for the subject assessments is the focus of this study guide but equally important is *how* you study. You can increase your chances of truly mastering the information by taking some simple, but effective steps.

Study Tips:

1. **Some foods aid the learning process.** Foods such as milk, nuts, seeds, rice, and oats help your study efforts by releasing natural memory enhancers called CCKs (*cholecystokinin*) composed of *tryptophan*, *choline*, and *phenylalanine*. All of these chemicals enhance the neurotransmitters associated with memory. Before studying, try a light, protein- rich meal of eggs, turkey, and fish. All of these foods release the memory enhancing chemicals. The better the connections, the more you comprehend.

Likewise, before you take a test, stick to a light snack of energy boosting and relaxing foods. A glass of milk, a piece of fruit, or some peanuts all release various memory-boosting chemicals and help you to relax and focus on the subject at hand.

2. **Learn to take great notes.** A by-product of our modern culture is that we have grown accustomed to getting our information in short doses (i.e. TV news sound bites or USA Today style newspaper articles.)

Consequently, we've subconsciously trained ourselves to assimilate information better in neat little packages. If your notes are scrawled all over the paper, it fragments the flow of the information. Strive for clarity. Newspapers use a standard format to achieve clarity. Your notes can be much clearer through use of proper formatting. A very effective format is called the "Cornell Method."

Take a sheet of loose- leaf lined notebook paper and draw a line all the way down the paper about 1-2" from the left-hand edge.

Draw another line across the width of the paper about 1-2" up from the bottom. Repeat this process on the reverse side of the page.

Look at the highly effective result. You have ample room for notes, a left hand margin for special emphasis items or inserting supplementary data from the textbook, a large area at the bottom for a brief summary, and a little rectangular space for just about anything you want.

3. Get the concept then the details. Too often we focus on the details and don't gather an understanding of the concept. However, if you simply memorize only dates, places, or names, you may well miss the whole point of the subject.

A key way to understand things is to put them in your own words. If you are working from a textbook, automatically summarize each paragraph in your mind. If you are outlining text, don't simply copy the author's words.

Rephrase them in your own words. You remember your own thoughts and words much better than someone else's, and subconsciously tend to associate the important details to the core concepts.

4. Ask Why? Pull apart written material paragraph by paragraph and don't forget the captions under the illustrations.

Example: If the heading is "Stream Erosion", flip it around to read "Why do streams erode?" Then answer the questions.

If you train your mind to think in a series of questions and answers, not only will you learn more, but it also helps to lessen the test anxiety because you are used to answering questions.

5. Read for reinforcement and future needs. Even if you only have 10 minutes, put your notes or a book in your hand. Your mind is similar to a computer; you have to input data in order to have it processed. *By reading, you are creating the neural connections for future retrieval.* The more times you read something, the more you reinforce the learning of ideas.

Even if you don't fully understand something on the first pass, *your mind stores much of the material for later recall.*

6. Relax to learn so go into exile. Our bodies respond to an inner clock called biorhythms. Burning the midnight oil works well for some people, but not everyone.

If possible, set aside a particular place to study that is free of distractions. Shut off the television, cell phone, pager and exile your friends and family during your study period.

If you really are bothered by silence, try background music. Light classical music at a low volume has been shown to aid in concentration over other types. Music that evokes pleasant emotions without lyrics are highly suggested. Try just about anything by Mozart. It relaxes you.

7. **Use arrows not highlighters.** At best, it's difficult to read a page full of yellow, pink, blue, and green streaks. Try staring at a neon sign for a while and you'll soon see that the horde of colors obscure the message.

A quick note, a brief dash of color, an underline, and an arrow pointing to a particular passage is much clearer than a horde of highlighted words.

8. **Budget your study time.** Although you shouldn't ignore any of the material, *allocate your available study time in the same ratio that topics may appear on the test.*

Testing Tips:

1. **Get smart, play dumb. Don't read anything into the question.** Don't make an assumption that the test writer is looking for something else than what is asked. Stick to the question as written and don't read extra things into it.

2. **Read the question and all the choices twice before answering the question.** You may miss something by not carefully reading, and then re-reading both the question and the answers.

If you really don't have a clue as to the right answer, leave it blank on the first time through. Go on to the other questions, as they may provide a clue as to how to answer the skipped questions.

If later on, you still can't answer the skipped ones . . . **Guess.** The only penalty for guessing is that you *might* get it wrong. Only one thing is certain; if you don't put anything down, you will get it wrong!

3. **Turn the question into a statement.** Look at the way the questions are worded. The syntax of the question usually provides a clue. Does it seem more familiar as a statement rather than as a question? Does it sound strange?

By turning a question into a statement, you may be able to spot if an answer sounds right, and it may also trigger memories of material you have read.

4. **Look for hidden clues.** It's actually very difficult to compose multiple-choice questions without giving away part of the answer in the options presented.

In most multiple-choice questions you can often readily eliminate one or two of the potential answers. This leaves you with only two real possibilities and automatically your odds go to Fifty-Fifty for very little work.

5. **Trust your instincts.** For every fact that you have read, you subconsciously retain something of that knowledge. On questions that you aren't really certain about, go with your basic instincts. **Your first impression on how to answer a question is usually correct.**

6. **Mark your answers directly on the test booklet.** Don't bother trying to fill in the optical scan sheet on the first pass through the test.

Just be very careful not to miss-mark your answers when you eventually transcribe them to the scan sheet.

7. **Watch the clock!** You have a set amount of time to answer the questions. Don't get bogged down trying to answer a single question at the expense of 10 questions you can more readily answer.

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
	IA	IIA	IIIB	IVB	VB	VIB	VIIA	VIIIB	VIIIA	IX	X	XI	XII	IIIA	IVA	VIA	VIIA	VIIIA	
1	hydrogen 1 H 1.0079	beryllium 4 Be 9.0122	lithium 3 Li 6.941	boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	neon 10 Ne 20.180	11	12	13	14	15	16	17	18	helium 2 He 4.0026	
2	sodium 11 Na 22.990	magnesium 12 Mg 24.305	calcium 20 Ca 40.078	potassium 19 K 39.098	scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	zinc 30 Zn 65.409	gallium 31 Ga 69.723	germanium 32 Ge 72.64	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.798
3	rubidium 37 Rb 85.468	strontium 38 Sr 87.62	yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	tantalum 72 Ta 180.95	vanadium 51 V 50.942	chromium 52 Cr 51.996	manganese 53 Mn 54.938	iron 54 Fe 55.845	cobalt 55 Co 58.933	nickel 56 Ni 58.693	zinc 64 Zn 65.409	gallium 69 Ga 69.723	germanium 72 Ge 72.64	arsenic 75 As 74.922	selenium 78 Se 78.96	bromine 80 Br 79.904	krypton 84 Kr 83.798
4	cesium 55 Cs 132.91	barium 56 Ba 137.33	lanthanum 57 La 138.91	hafnium 72 Hf 178.49	tantalum 73 Ta 180.95	vanadium 51 V 50.942	chromium 52 Cr 51.996	manganese 53 Mn 54.938	iron 54 Fe 55.845	cobalt 55 Co 58.933	nickel 56 Ni 58.693	zinc 64 Zn 65.409	gallium 69 Ga 69.723	germanium 72 Ge 72.64	arsenic 75 As 74.922	selenium 78 Se 78.96	bromine 80 Br 79.904	krypton 84 Kr 83.798	rubidium 87 Rb 85.468
5	francium 87 Fr [223]	radium 88 Ra [226]	actinium 89 Ac [227]	thorium 90 Th 232.04	protactinium 91 Pa 231.04	uranium 92 U 238.03	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]	lawrencium 103 Lr [262]	104	105	106
6	109	110	111	112	113	114	115	116	117	118	119	120	121	122	123	124	125	126	127
7	128	129	130	131	132	133	134	135	136	137	138	139	140	141	142	143	144	145	146

Periodic Table of the Elements

*Lanthanoids

**Actinoids

Atomic mass values from IUPAC review (2001): http://www.iupac.org/reports/periodic_table/

Part I Understanding and Skill in Chemistry

This book follows the California numbering system for labeling all chemistry Subject Matter Requirement (SMR) Domains. Domains 1 through 7 in this section are referred to as Chemistry Domains. Domains 11 and 12 in the following section are referred to as General Chemistry Domains enumerated within the broader scope of general science.

Domain 1.0 Atomic and Molecular Structure

Competency 1.1 Periodic Table and Periodicity

What the ocean was to the child, the Periodic Table is to the chemist.

—K. Barry Sharpless (Nobel prize in Chemistry, 2001)

Skill 1.1a- Differentiate periodic groups and families of elements and their properties

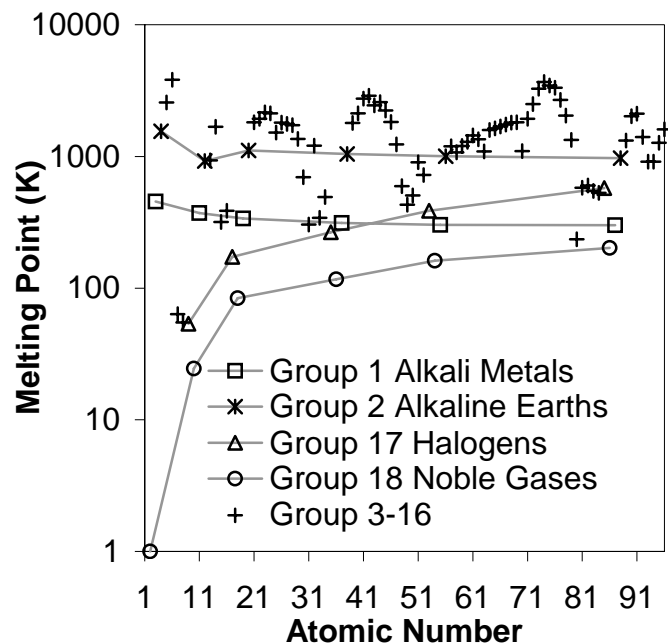
The construction and organization of the periodic table are described in **Skill 9.1k**.

Groups 1, 2, 17, and 18 are often identified with the group names shown on the table to the right. Groups 3 through 12 are called the **transition metals**. The lanthanoid series is contained in period 6, and the actinoid series is in period 7. The two series together are called the **inner transition metals**. The locations of the transition and inner transition metals in the periodic table are discussed further in **Skill 1.1b**. Elements in the periodic table are also divided into broad categories of metals, nonmetals, and semimetals as discussed in **Skill 9.1l**.

Alkali metals
 Halogens
 Alkaline earth metals
 Noble gases

Several elements are found as **diatomic molecules**: (H_2 , N_2 , O_2 , and the halogens: F_2 , Cl_2 , Br_2 , and I_2). Mnemonic devices to remember the diatomic elements are: “ $Br_2I_2N_2Cl_2H_2O_2F_2$ ” (pronounced “Brinklehof”) and “**Have No Fear Of Ice Cold Beer.**” These molecules are attracted to one another using **weak London dispersion forces** (see **Skill 1.3d**).

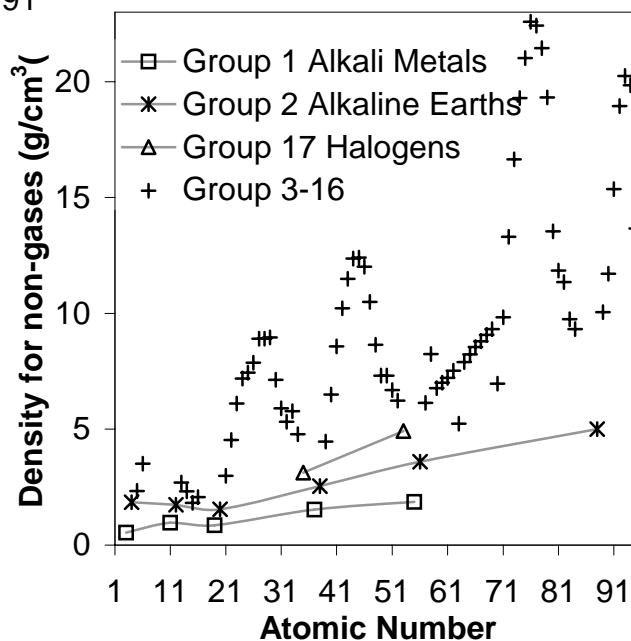
Note that **hydrogen** is not an alkali metal. Hydrogen is a colorless gas and is the most abundant element in the universe, but H_2 is very rare in the atmosphere because it is light enough to escape gravity and reach outer space. Hydrogen atoms form more compounds than any other element.



Alkaline earth metals (group 2 elements) are grey, metallic solids. They are harder, denser, and have a higher melting point than the alkali metals (see asterisk data points in the figures), but values for these properties are still low compared to most of the transition metals. Measures of metallic bond strength like melting points for alkaline earths do not follow a simple trend down the periodic table.

When cut by a knife, the exposed surface of an **alkali metal or alkaline earth metal** quickly turns into an oxide. These elements **do not occur in nature as free metals**. Instead, they react with many other elements to form white or grey water-soluble salts. With some exceptions, the oxides of group 1 elements have the formula M_2O , their hydrides are MH , and their halides are MX (for example, $NaCl$). The oxides of group 2 elements have the formula MO , their hydrides are MH_2 , and their halides are MX_2 .

Alkali metals are shiny, soft, metallic solids. They have **low melting points and low densities** compared with other metals because they have a weaker metallic bond (see the square data points in the figures to the left and below). Measures of intermolecular attractions including their **melting points decrease further down the periodic table due to weaker metallic bonds** as the size of atoms increases. See **Skill 1.3d** for a discussion of metallic bonding. Most salts with an alkali metal cation are always soluble (see **Skill 4.1a**).



Copper, silver, and gold (group 11) are known as the **noble metals** or **coinage metals** because they are very unreactive.

Halogens (group 17 elements) have an irritating odor. Unlike the metallic bonds between alkali metals, weak **London forces between halogen molecules increase in strength further down the periodic table**, increasing their melting points as shown by the triangular data points in the figures on the previous page. Weak London forces (see **Skill 1.3d**) make Br_2 a liquid and I_2 a solid at 25 °C. The lighter halogens are gases. Halogens form a wide variety of oxides and also combine with other halogens. They combine with hydrogen to form HX gases, and these compounds are also commonly used as acids (hydrofluoric, hydrochloric, etc.) in aqueous solution. Halogens form salts with metals by gaining electrons to become X^- ions. Halogen compounds are called halides. Astatine is an exception to many of these properties because it is an artificial metalloid.

Noble gases (group 18 elements) have no color or odor and exist as **individual gas atoms** that experience London forces. These attractions also increase with period number as shown by the circular data points in the figures on the previous page. Noble gases are **nearly chemically inert**. The heavier noble gases form a number of compounds with oxygen and fluorine such as KrF_2 and XeO_4

Skill 1.1b- Relate valence electrons and the electron shell structures to an element's position in the periodic table

The position of an element in the periodic table may be related to its electron configuration, and this configuration in turn results from the quantum theory describing the filling of a shell of electrons. In this skill, we will take this theory as our starting point. However, it should be remembered that it is the correlation with properties—not with electron arrangements—that have placed the periodic table at the beginning of most chemistry texts.

Quantum numbers

The quantum-mechanical solutions from the Schrödinger Equation (see **Skill 1.2a**) utilize three quantum numbers (n , l , and m_l) to describe an orbital and a fourth (m_s) to describe an electron in an orbital. This model is useful for understanding the frequencies of radiation emitted and absorbed by atoms and chemical properties of atoms.

The **principal quantum number n** may have positive integer values (1, 2, 3, ...). n is a measure of the **distance** of an orbital from the nucleus, and orbitals with the same value of n are said to be in the same **shell**. This is analogous to the Bohr model of the atom (see **Skill 1.2a**). Each shell may contain up to $2n^2$ electrons.

The **azimuthal quantum number l** may have integer values from 0 to $n-1$. l describes the angular momentum of an orbital. This determines the orbital's **shape**. Orbitals with the same value of n and l are in the same **subshell**, and each subshell may contain up to $4l + 2$ electrons. Subshells are usually referred to by the principle quantum number followed by a letter corresponding to l as shown in the following table:

Azimuthal quantum number l	0	1	2	3	4
Subshell designation	s	p	d	f	g

The **magnetic quantum number m_l or m** may have integer values from $-l$ to l . m_l is a measure of how an individual orbital responds to an external magnetic field, and it often describes an orbital's **orientation**. A subscript—either the value of m_l or a function of the x -, y -, and z -axes—is used to designate a specific orbital. See **Skill 1.2a** for images of electron density regions for a few orbitals of hydrogen. $n=3$, $l=2$, and $m_l=0$ for the $3d_0$ orbital. Each orbital may hold up to two electrons.

The **spin quantum number m_s or s** has one of two possible values: $-1/2$ or $+1/2$. m_s differentiates between the two possible electrons occupying an orbital. Electrons moving through a magnet behave as if they were tiny magnets themselves spinning on their axis in either a clockwise or counterclockwise direction. These two spins may be described as $m_s = -1/2$ and $+1/2$ or as down and up.

The **Pauli exclusion principle** states that **no two electrons in an atom may have the same set of four quantum numbers**.

The following table summarizes the relationship among n , l , and m_l through $n=3$:

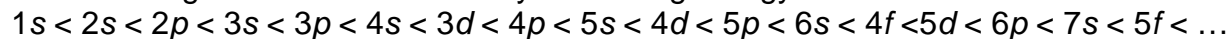
n	l	Subshell	m_l	Orbitals in subshell	Maximum number of electrons in subshell
1	0	1s	0	1	2
2	0	2s	0	1	2
	1	2p	-1, 0, 1	3	6
3	0	3s	0	1	2
	1	3p	-1, 0, 1	3	6
	2	3d	-2, -1, 0, 1, 2	5	10

Subshell energy levels

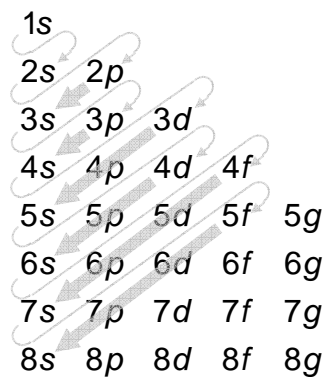
In single- electron atoms (H, He^+ , and Li^{2+}) above the ground state, subshells within a shell are all at the same energy level, and an orbital's energy level is only determined by n . However, in all other atoms, multiple electrons repel each other. Electrons in orbitals closer to the nucleus create a screening or **shielding effect** on electrons further away from the nucleus, preventing them from receiving the full attractive force of the nucleus.

In multi- electron atoms, both n and l determine the energy level of an orbital. In the absence of a magnetic field, **orbitals in the same subshell with different m_l all have the same energy** and are said to be **degenerate orbitals**.

The following list orders subshells by increasing energy level:



This list may be constructed by arranging the subshells according to n and l and drawing diagonal arrows as shown below:

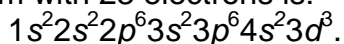


Drawing electron shell structures

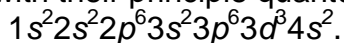
Electron shell structures (also called electron arrangements) in an atom may be represented using three methods: an **electron configuration**, an **orbital diagram**, or an **energy level diagram**.

All three methods require knowledge of the subshells occupied by electrons in a certain atom. The **Aufbau principle** or **building-up rule** states that **electrons at ground state fill orbitals starting at the lowest available energy levels**.

An **electron configuration** is a **list of subshells** with superscripts representing the **number of electrons** in each subshell. For example, an atom of boron has 5 electrons. According to the Aufbau principle, two will fill the 1s subshell, two will fill the higher energy 2s subshell, and one will occupy the 2p subshell which has an even higher energy. The electron configuration of boron is $1s^2 2s^2 2p^1$. Similarly, the electron configuration of a vanadium atom with 23 electrons is:



Configurations are also written with their principle quantum numbers together:

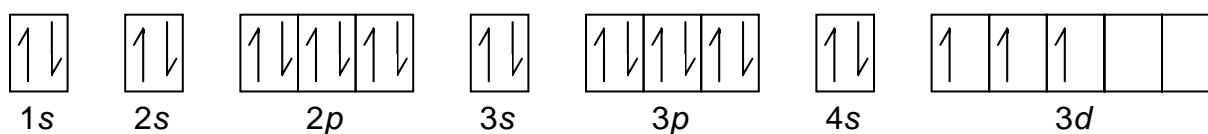


Electron configurations are often written to emphasize the outermost electrons. This is done by writing the symbol in brackets for the element with a full p subshell from the previous shell and adding the **outer electron configuration** onto that configuration. The element with the last full p subshell will always be a noble gas from the right-most column of the periodic table (see **Skill 1.1a** and **Skill 9.11**). For the vanadium example, the element with the last full p subshell has the configuration $1s^2 2s^2 2p^6 3s^2 3p^6$. This is ${}_{18}\text{Ar}$. The configuration of vanadium may then be written as $[\text{Ar}]4s^2 3d^3$ where $4s^2 3d^3$ is the outer electron configuration.

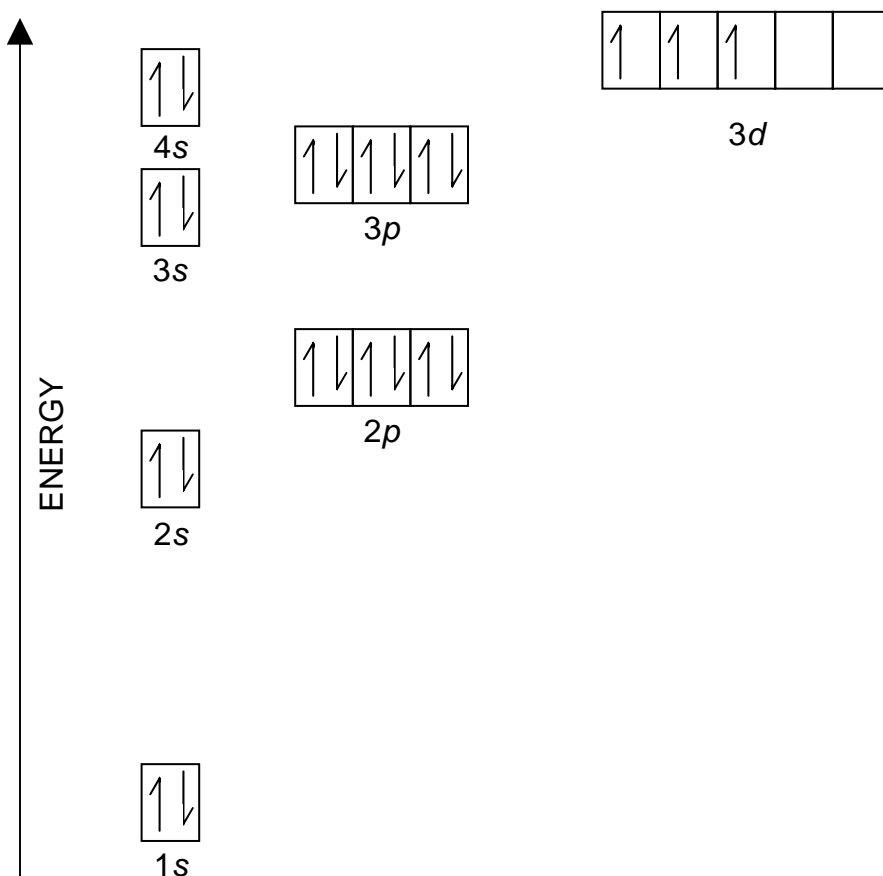
Electron shell structures may also be written by noting the number of electrons in each shell. For vanadium, this would be:

2, 8, 11, 2.

Orbital diagrams assign electrons to individual orbitals so the energy state of individual electrons may be found. This requires knowledge of how electrons occupy orbitals within a subshell. **Hund's rule** states that **before any two electrons occupy the same orbital, other orbitals in that subshell must first contain one electron each with parallel spins**. Electrons with up and down spins are shown by half- arrows, and these are placed in lines of orbitals (represented as boxes or dashes) according to Hund's rule, the Aufbau principle, and the Pauli exclusion principle. Below is the orbital diagram for vanadium:



An **energy level diagram** is an orbital diagram that shows subshells with higher energy levels higher up on the page. The energy level diagram of vanadium is:

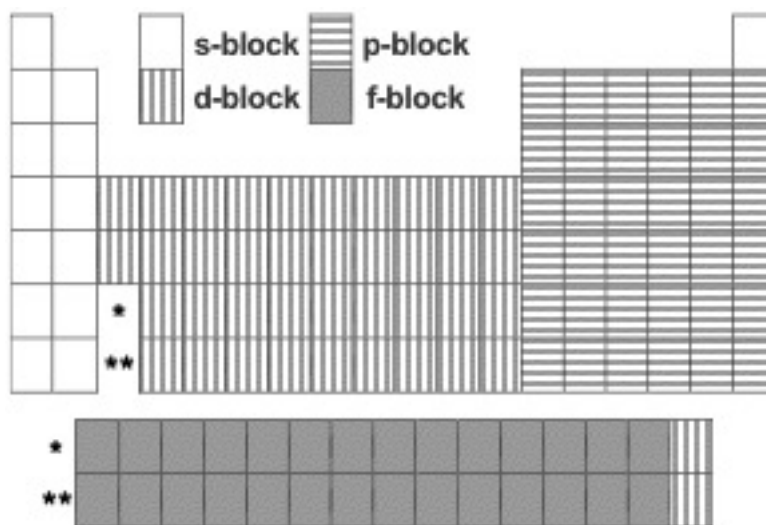


Valence shell electrons and the periodic table

Electrons in the **outermost shell** are called **valence shell electrons**. For example, the electron configuration of Se is $[\text{Ar}]4s^23d^{10}4p^4$, and its valence shell electron configuration is $4s^24p^4$.

The **periodic table** may be related to the electron shell structure of any element. The table may be divided up into **blocks corresponding to the subshell** designation of the most recent orbital to be filled by the building-up rule.

Elements in the s- and p-blocks are known as **main-group elements**. The d-block elements are called **transition metals**. The f-block elements are called **inner transition metals**.



The maximum number of electrons in each subshell (2, 6, 10, or 14) determines the number of elements in each block, and the order of energy

levels for subshells create the pattern of blocks. These blocks also usually correspond to the value of l for the **outermost electron** of the atom. This has important consequences for the physical and chemical properties of the elements as shown in **Skill 1.1c**. The outermost shell or valence shell principle quantum number (for example, 4 for Se) is also the period number for the element in the table.

Atoms in the d- and f- blocks often have unexpected electron shell structures that cannot be explained using simple rules. Some heavy atoms have unknown electron configurations because the number of different frequencies of radiation emitted and absorbed by these atoms is very large.

http://www.cowtownproductions.com/cowtown/genchem/08_07T1.htm contains a brief tutorial on energy level diagrams.

<http://www.colorado.edu/physics/2000/applets/a2.html> contains (among other things) energy level diagrams and animations of electron shells and nuclei.

<http://intro.chem.okstate.edu/WorkshopFolder/Electronconfnew.html> animates the building up of energy level diagrams.

Skill 1.1c- Predict periodic trends including electronegativity, ionization energy, and the relative sizes of ions and atoms

Electronegativity

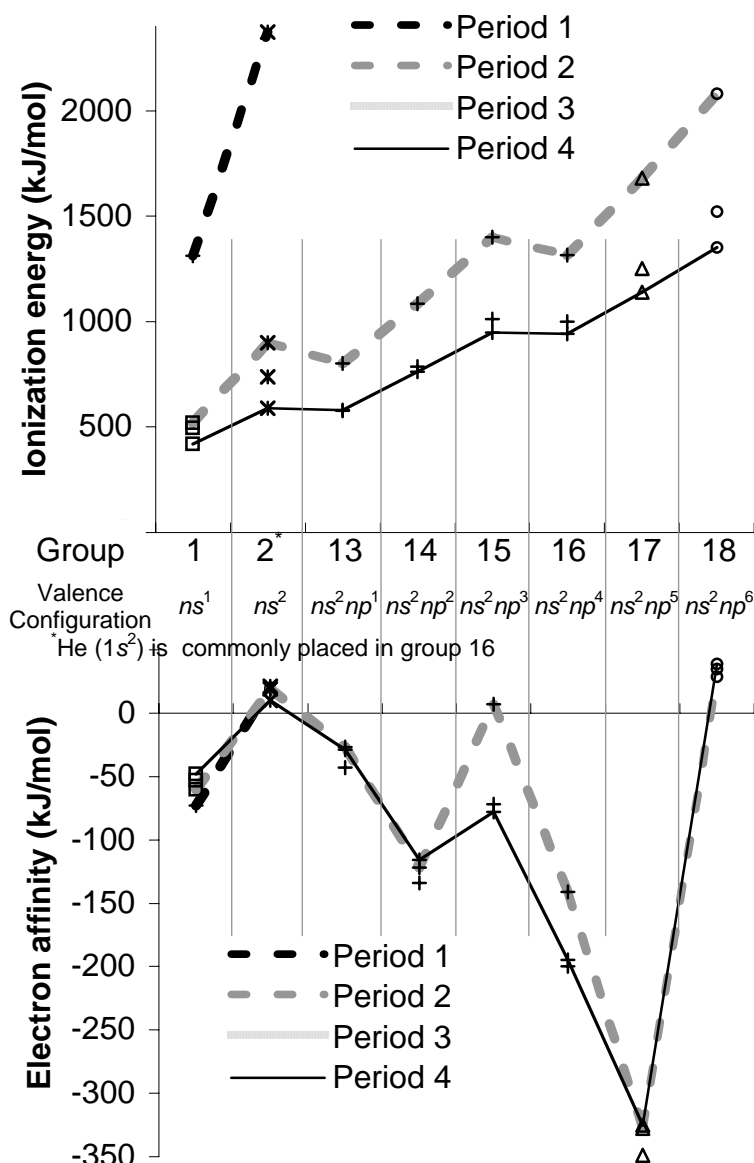
Electronegativity measures the ability of an atom to attract electrons in a chemical bond. The most metallic elements (see **Skill 9.1I**) at the lower left of the periodic table have the lowest electronegativity. The most nonmetallic have the highest electronegativity. The impact of electronegativity on chemical bonding is discussed in **Skill 1.3a**.

Physics of electrons and stability of electron configurations

For an isolated atom, the **most stable system of valence electrons is a filled set of orbitals** (see **Skill 1.1b** and **Skill 9.1I**). For the main group elements, this corresponds to group 18 (ns^2np^6 and $1s^2$ for helium), and, to a lesser extent, group 2 (ns^2). The next most stable state is a set of degenerate half-filled orbitals. These occur in group 15 (ns^2np^3). The least stable valence electron configuration is a single electron with no other electrons in similar orbitals. This occurs in group 1 (ns^1) and to a lesser extent in group 13 (ns^2np^1).

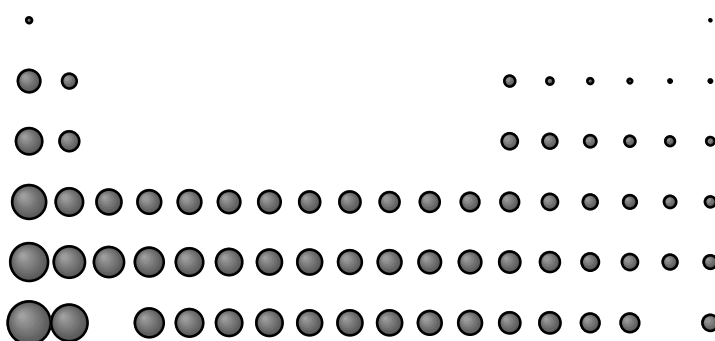
An atom's first **ionization energy** is the energy required to remove one electron by the reaction $M(g) \rightarrow M^+(g) + e^-$. Periodicity is in the opposite direction from the trend for atomic radius. The most metallic atoms have electrons further from the nucleus, and these are easier to remove.

An atom's **electron affinity** is the energy released when one electron is added by the reaction $M(g) + e^- \rightarrow M^-(g)$.



A large negative number for the exothermic reaction indicates a high electron affinity. Halogens have the highest electron affinities.

Trends in **ionization energy and electron affinity** within a period reflect the **stability of valence electron configurations**. A stable system requires more energy to change and releases less when changed. Note the peaks in stability for groups 2, 13, and 16 to the right.



The **size of an atom** is not an exact distance due to of the probabilistic nature of electron density (see Skill 1.2a), but we may compare radii among different atoms using some standard. As seen to the right, the sizes of neutral atoms increase

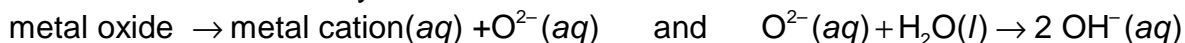
with period number and decrease with group number. As period number increases, the outermost electrons are shielded from the positive charge of the nucleus by more electrons in shells further in, so they are held less tightly. As group number increases, electrons are added to the same shell, so they experience little additional shielding, but the increased number of protons holds them in more closely. This trend is similar to the trend for metallic character. The smallest atom is helium.

The **size of an ion** is also not an exact distance due to the probabilistic nature of electron density, but different ion sizes may be compared. The size of different ions with the same number of electrons decreases as the number of protons increases because more protons provide a greater attractive force (e.g., the size of $O^{2-} > F^- > Ne > Na^+ > Mg^{2+} > Al^{3+}$ for ions with 10 electrons). Cations are smaller than the same parent atom ($Na^+ < Na$) because of decreased repulsion among electrons and anions are larger than the same parent atom ($Cl^- > Cl$) because of increased electron repulsion. Ions of the same charge show periodic trends identical to the trends for neutral atoms. Sizes increase with period number ($F^- < Cl^- < Br^- < I^-$) and decrease with group number ($Na^+ > Mg^+ > Al^+$).

Periodic trends in **melting point** and **density** were charted in **Skill 1.1a**.

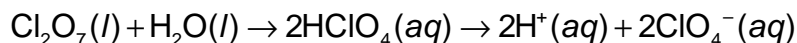
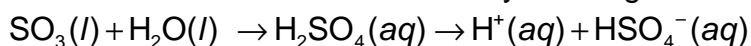
Intermolecular forces contribute to density by bringing nuclei closer to each other, so the periodicity in density is similar to trends for melting point. However, group-to-group differences are superimposed on a general trend for **density to increase with period number** because heavier nuclei make the material denser. See **Skill 1.3d** for other properties altered by intermolecular forces.

Metal oxides form basic solutions in water because the ionic bonds break apart and the O^{2-} ion reacts to form hydroxide ions:



Ionic oxides containing a large cation with a low charge (Rb_2O , for example) are most soluble and form the strongest bases.

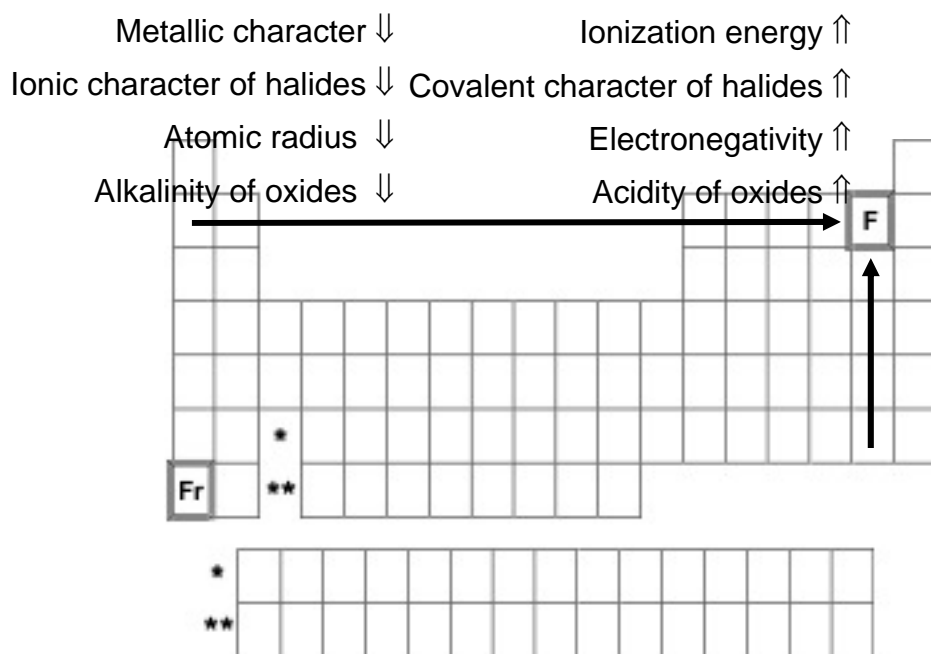
Covalent oxides form acidic solutions in water by reacting with water. For example:



Covalent oxides at high oxidation states and high electronegativities form the strongest acids. Acids and bases are discussed in **Competency 4.2**. For this skill, note that the periodic trends for acid and base strength of the oxide of an element follows the same pattern we've seen before.

Summary

A summary of periodic trends is shown to the right. The properties tend to decrease or increase as shown depending on a given element's proximity to fluorine in the table.



<http://jcrystal.com/steffenweber/JAVA/jpt/jpt.html> contains an applet of the periodic table and trends.

<http://www.webelements.com> is an on-line reference for information on the elements.

<http://www.uky.edu/Projects/Chemcomics/> has comic book pages for each element.

Competency 1.2 Atomic Structure

Skill 1.2a- Analyze the evolution of the atomic model

Dalton

The existence of fundamental units of matter called atoms of different types called elements was proposed by ancient philosophers without any evidence to support the belief. Modern atomic theory is credited to the work of **John Dalton** published in 1803-1807. Observations made by him and others about the composition, properties, and reactions of many compounds led him to develop the following postulates:

- 1) Each element is composed of small particles called atoms.
- 2) All atoms of a given element are identical in mass and other properties.
- 3) Atoms of different elements have different masses and differ in other properties.
- 4) Atoms of an element are not created, destroyed, or changed into a different type of atom by chemical reactions.
- 5) Compounds form when atoms of more than one element combine.
- 6) In a given compound, the relative number and kind of atoms are constant.

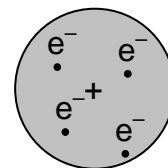
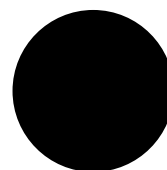
Dalton determined and published the known relative masses of a number of different atoms. He also formulated the law of partial pressures (see **Skill 3.1f**). Dalton's work focused on the ability of atoms to arrange themselves into molecules and to rearrange themselves via chemical reactions, but he did not investigate the composition of atoms themselves. **Dalton's model of the atom** was a tiny, indivisible, indestructible **particle** of a certain mass, size, and chemical behavior, but Dalton did not deny the possibility that atoms might have a substructure.

ELEMENTS			
	Hydrogen	1	1
	Azote	5	5
	Carbon	5	5
	Oxygen	7	7
	Phosphorus	9	9
	Sulphur	13	13
	Magnesia	20	20
	Lime	24	24
	Soda	28	28
	Potash	47	47
	Strontian	86	86
	Barries	88	88
	Iron	50	50
	Zinc	56	56
	Copper	56	56
	Lead	90	90
	Silver	190	190
	Gold	190	190
	Platina	190	190
	Mercury	197	197

Dalton's table of atomic symbols and masses

Thomson

Joseph John Thomson, often known as **J. J. Thomson**, was the first to examine this substructure. In the mid- 1800s, scientists had studied a form of radiation called "cathode rays" or "electrons" that originated from the negative electrode (cathode) when electrical current was forced through an evacuated tube. Thomson determined in 1897 that **electrons have mass**, and because many different cathode materials release electrons, Thomson proposed that the **electron is a subatomic particle**. **Thomson's model of the atom** was a uniformly positive particle with electrons contained in the interior.



This has been called the "plum- pudding" model of the atom where the pudding represents the uniform sphere of positive electricity and the bits of plum represent electrons. For more on Thomson, see <http://www.aip.org/history/electron/jjhome.htm>.

Planck

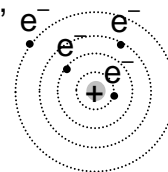
Max Planck determined in 1900 that **energy is transferred by radiation in exact multiples of a discrete unit of energy called a quantum**. Quanta of energy are extremely small, and may be found from the frequency of the radiation, ν , using the equation:

$$\Delta E = h\nu, 2h\nu, 3h\nu, \dots$$

where h is Planck's constant and $h\nu$ is a quantum of energy.

Rutherford

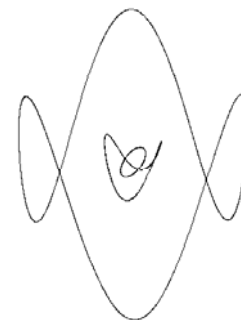
Ernest Rutherford studied atomic structure in 1910-1911 by firing a beam of alpha particles (see **Skill 7.1b**) at thin layers of gold leaf. According to Thomson's model, the path of an alpha particle should be deflected only slightly if it struck an atom, but Rutherford observed some alpha particles bouncing almost backwards, suggesting that **nearly all the mass of an atom is contained in a small positively charged nucleus**. **Rutherford's model of the atom** was an analogy to the sun and the planets. A small positively charged nucleus is surrounded by circling electrons and mostly by empty space. Rutherford's experiment is explained in greater detail in this flash animation:



<http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/ruther14.swf>.

De Broglie

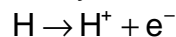
Depending on the experiment, radiation appears to have wave- like or particle- like traits. In 1923-1924, Louis de Broglie applied this **wave/particle duality to all matter with momentum**. The discrete distances from the nucleus described by Bohr corresponded to permissible distances where standing waves could exist. **De Broglie's model of the atom** described electrons as **matter waves in standing wave orbits** around the nucleus. The first three standing waves corresponding to the first three discrete distances are shown in the figure. An applet of de Broglie's model may be found here: <http://artsci-cwin.concordia.ca/facstaff/a-c/bird/c241/D1-part2.html>.



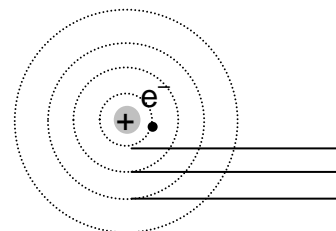
Bohr

Niels Bohr incorporated Planck's quantum concept into Rutherford's model of the atom in 1913 to explain the **discrete frequencies of radiation emitted and absorbed by atoms with one electron** (H, He⁺, and Li²⁺). This electron is attracted to the positive nucleus and is closest to the nucleus at the **ground state** of the atom. When the electron absorbs energy, it moves into an orbit further from the nucleus and the atom is

said to be in an electronically **excited state**. If sufficient energy is absorbed, the electron separates from the nucleus entirely, and the atom is ionized:



The energy required for ionization from the ground state is called the atom's **ionization energy**. The discrete frequencies of radiation emitted and absorbed by the atom correspond (using Planck's constant) to discrete energies and in turn to discrete distances from the nucleus. **Bohr's model of the atom** was a small positively charged nucleus surrounded mostly by empty space and by electrons orbiting at certain discrete distances ("shells") corresponding to discrete energy levels. Animations utilizing the Bohr model may be found at the following two URLs: <http://artsci-ccwin.concordia.ca/facstaff/a-c/bird/c241/D1.html> and <http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/linesp16.swf>.



Heisenberg

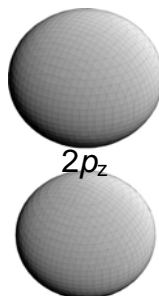
The realization that both matter and radiation interact as waves led Werner Heisenberg to the conclusion in 1927 that the act of observation and measurement requires the interaction of one wave with another, resulting in an **inherent uncertainty** in the location and momentum of particles. This inability to measure phenomena at the subatomic level is known as the **Heisenberg uncertainty principle**. A discussion of the principle and Heisenberg's other contributions to quantum theory is located here: <http://www.aip.org/history/heisenberg/>.

Schrödinger

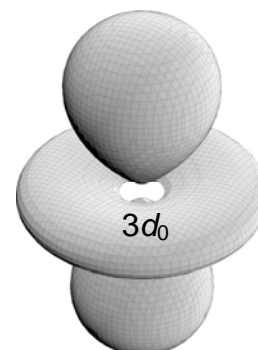
When Erwin Schrödinger studied the atom in 1925, he replaced the idea of precise orbits with regions in space called **orbitals** where electrons were likely to be found. **The Schrödinger equation** describes the **probability** that an electron will be in a given region of space, a quantity known as **electron density** or ψ^2 . The diagrams below are surfaces of constant ψ^2 found by solving the Schrödinger equation for the hydrogen atom 1s, $2p_z$ and $3d_0$ orbitals (see **Skill 1.1b**). Additional representations of solutions may be found here: <http://library.wolfram.com/webMathematica/Physics/Hydrogen.jsp>. **Schrödinger's model of the atom** is a mathematical formulation of quantum mechanics that describes the electron density of orbitals. It is the atomic model that has been in use from shortly after it was introduced up to the present.



CHEMISTRY



13



Skill 1.2b- Relate atomic spectroscopy and the photoelectric effect to the quantum structure of the atom

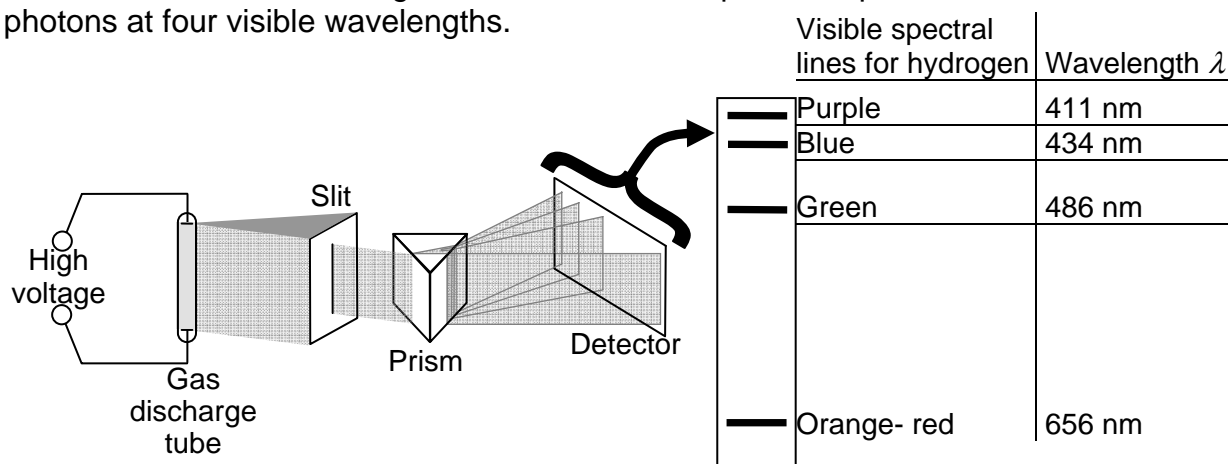
The quantum structure of the atom describes electrons in discrete energy levels surrounding the nucleus (see **Skill 1.2a**). When an electron moves from a high energy orbital to a lower energy orbital, a quantum of electromagnetic radiation is emitted, and for an electron to move from a low energy to a higher energy level, a quantum of radiation must be absorbed. The particle that carries this electromagnetic force is called a **photon**. The quantum structure of the atom predicts that only photons corresponding to certain wavelengths of light will be emitted or absorbed by atoms. These distinct wavelengths are measured by **atomic spectroscopy**.

In **atomic absorption spectroscopy**, a continuous spectrum (light consisting of all wavelengths) is passed through the element. The frequencies of absorbed photons are then determined as the electrons increase in energy. An **absorption spectrum** in the visible region usually appears as a rainbow of color stretching from red to violet interrupted by a few black lines corresponding to distinct wavelengths of absorption.

In **atomic emission spectroscopy**, the electrons of an element are excited by heating or by an electric discharge. The frequencies of emitted photons are then determined as the electrons release energy.

An **emission spectrum** in the visible region typically consists of lines of light at certain colors corresponding to distinct wavelengths of emission. The bands of emitted or absorbed light at these wavelengths are called **spectral lines**. **Each element has a unique line spectrum**. Light from a star (including the sun) may be analyzed to determine what elements are present.

A simple optical spectroscope separates visible light into distinct wavelengths by passing the light through a prism or diffraction grating. When electrons in hydrogen gas are excited inside a discharge tube, the emission spectroscope shown below detects photons at four visible wavelengths.



Quantum #	Radius
$n \rightarrow \infty$	$r_{\infty} \rightarrow \infty$
\vdots	\vdots
$n = 5$	$r_5 = 25a_0$
$n = 4$	$r_4 = 16a_0$
$n = 3$	$r_3 = 9a_0$
$n = 2$	$r_2 = 4a_0$
$n = 1$	$r_1 = a_0$
	\oplus (H nucleus)

An electron may exist at distinct radial distances (r_n) from the nucleus. These distances are proportional to the square of the **principal quantum number**, n . For a hydrogen atom (shown at left), the proportionality constant is called the **Bohr radius** ($a_0 = 5.29 \times 10^{-11}$ m). This value is the mean distance of an electron from the nucleus at the ground state of $n = 1$. The distances of other electron shells are found by the formula:

$$r_n = a_0 n^2.$$

As $n \rightarrow \infty$, the electron is no longer part of the hydrogen atom. Ionization occurs and the atom becomes an H^+ ion.

Photon wavelength (λ) in meters and frequency (ν) in reciprocal seconds are inversely proportional to each other. The proportionality constant between them is the **speed of light** ($c = 3.00 \times 10^8$ m/s):

$$\lambda = \frac{c}{\nu} \quad \text{and} \quad \nu = \frac{c}{\lambda}.$$