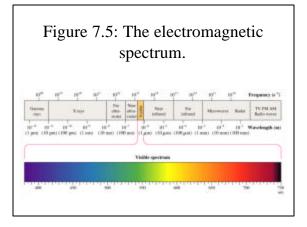
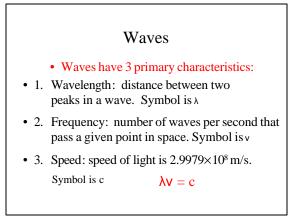
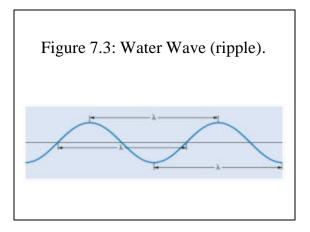


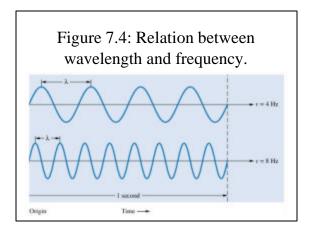
# Electromagnetic Radiation

• Radiant energy that exhibits wavelength-like behavior and travels through space at the speed of light in a vacuum.









Wavelength and frequency can be interconverted.

 $\mathbf{n} = c/\mathbf{l}$ 

$$n = frequency (s^{-1})$$

$$l = wavelength (m)$$

• 
$$c = \text{speed of light (m s^{-1})}$$

#### The Wave Nature of Light

• What is the wavelength of yellow light with a frequency of 5.09 x 10<sup>14</sup> s<sup>-1</sup>? (Note: s<sup>-1</sup>, commonly referred to as Hertz (Hz) is defined as "cycles or waves per second".)

– If  $c = v\lambda$ , then rearranging, we obtain  $\lambda = c/v$ 

$$\mathbf{l} = \frac{(3.00^{-10^8} \text{m/s})}{5.09^{-10^{14}} \text{s}^{-1}} = 5.89^{-10^{-7}} \text{m or } 589 \text{ nm}$$

# The Wave Nature of Light

• What is the frequency of violet light with a wavelength of 408 nm? (see Figure 7.5)

– If  $c = \nu \lambda$ , then rearranging, we obtain  $\nu = c/\lambda$ .

$$\mathbf{n} = \frac{3.00 \times 10^8 \, m/s}{408 \times 10^{-9} \, m} = 7.35 \times 10^{14} \, s^{-1}$$

# **Quantum Effects and Photons**

- By the early part of twentieth century, the wave theory of light seemed to be well entrenched.
  - In 1905, Albert Einstein proposed that light had both wave and particle properties as observed in the **photoelectric effect.** (see Figure 7.6)
  - Einstein based this idea on the work of a German physicist, Max Planck.

#### **B** Quantum Effects and Photons

#### 1 Planck's Quantization of Energy

In order to explain the observation that the ?'s of light emitted from a hot solid was dependent on the temperature of the solid, Planck observed that the atoms in the solid vibrated with a definite ?, depending on the solid, and that the atom could only have certain energies of vibration, E.

#### $\mathbf{E} = \mathbf{n} \mathbf{h}$ ?

where n is the quantum number, h is Planck's constant, and ? is the frequency of vibration. Thus, the energy of the atoms is emitted in "packets" or quantized.

#### Planck's Constant

Transfer of energy is quantized, and can only occur in discrete units, called quanta.

$$DE = hn = \frac{hc}{l}$$

- $\Delta E$  = change in energy, in J
- $h = \text{Planck's constant}, 6.626 \times 10^{-34} \text{ J s}$

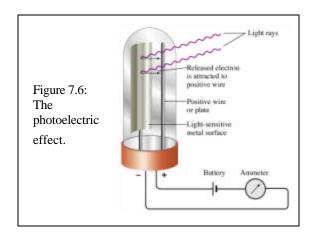
 $\mathbf{n} = \text{frequency, in s}^{-1}$ 

**l** = wavelength, in m

### **Quantum Effects and Photons**

#### Photoelectric Effect

- The photoelectric effect is the ejection of electrons from the surface of a metal when light shines on it. (see Figure 7.6)
- Electrons are ejected only if the light exceeds a certain "threshold" frequency.
- Violet light, for example, will cause potassium to eject electrons, but <u>no amount of red light</u> (which has a lower frequency) has <u>any</u> effect.



# **Quantum Effects and Photons**

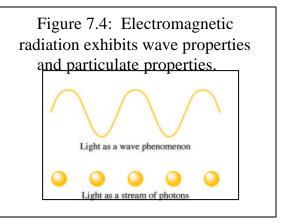
#### Photoelectric Effect

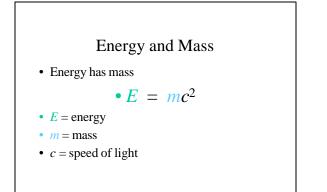
- Einstein's assumption that an electron is ejected when struck by a <u>single photon</u> implies that it behaves like a particle.
- When the photon hits the metal, its energy, hn is taken up by the electron.
- The photon ceases to exist as a particle; it is said to be "absorbed."

## **Quantum Effects and Photons**

#### Photoelectric Effect

- The "wave" and "particle" pictures of light should be regarded as complementary views of the same physical entity.
- This is called the **wave -particle duality** of light.
- The equation E = hv displays this duality; E is the energy of the "particle" photon, and v is the frequency of the associated "wave."



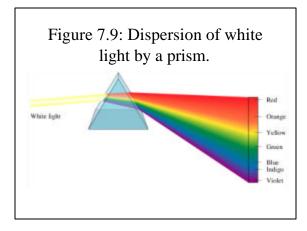


Energy and Mass  

$$E_{\text{photon}} = \frac{hc}{l}$$
  
 $m_{\text{photon}} = \frac{h}{lc}$   
(Hence the dual nature of light.)

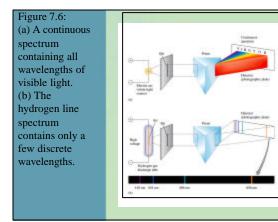
"They could but make the best of it, and went around with woebegone faces sadly complaining that on Mondays, Wednesdays, and Fridays they must look on light as a wave; on Tuesdays, Thursdays and Saturdays as a particle. On Sundays they simply prayed."

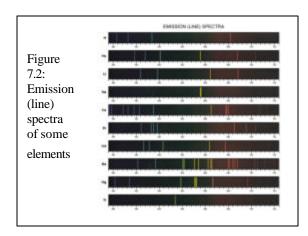
- Banesh Hoffman, The Strange Story of the Quantum, Dover Publications, Inc. New York 1959.



# Atomic Spectrum of Hydrogen

- Continuous spectrum: Contains all the wavelengths of light.
- Line (discrete) spectrum: Contains only some of the wavelengths of light.







If we pass light emitted from a heated gas, we obtain a line spectrum which contains only a few lines (or colors) or specific wavelengths characteristic of the gas. In 1885, J.J. Balmer showed that for the H atom, the observed visible ?'s were given by the equation:

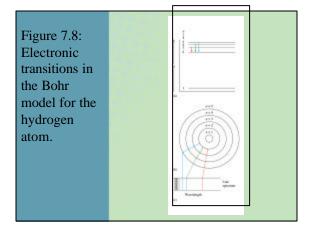
?' =  $1/? = 1.097 \times 10^7/m [(1/2^2)-(1/n^2)]$ Where n is an integer > 2 The constant, 1.097 x 10<sup>7</sup>, is called the Rydberg constant (R) and was determined experimentally.

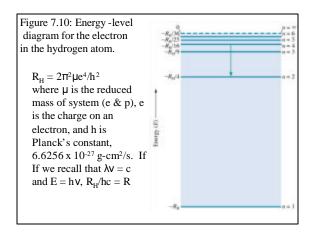
# The Bohr Model

• Ground State: The lowest possible energy state for an atom (n = 1).

- 2 Bohr's postulates developed to account for a) the stability of the H atom (e does not fall into the nucleus) and b) the observed line spectrum
  - a Energy Level Postulate an electron can have only *specific* energy values in an atom, called energy levels.

For the H atom:  $\mathbf{E} = -(\mathbf{R}_H/n^2)$  where  $\mathbf{R}_H$  is a constant with energy units and n is the principal quantum number

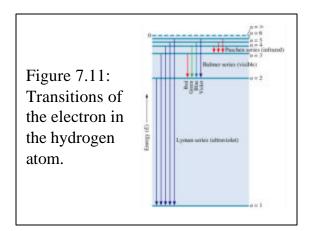


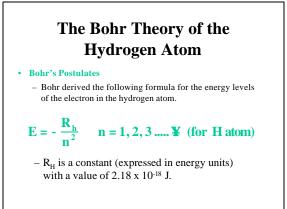


b Transitions between energy levels - an electron in an atom can change energy only by going from one energy level to another energy level. When it changes energy, it undergoes a transition and  $E_f$  -  $E_i = h$ ?

Bohr demonstrated that :

 $R_{\rm H}$  /hc = R = 1.097 x 10<sup>7</sup>/m from Balmer's formula. According to Bohr's theory, light is emitted from a H atom when its electrons drops from a higher to lower energy level. See Fig 7.11





# The Bohr Theory of the Hydrogen Atom

 Bohr's Postulates

 When an electron undergoes a transition from a higher energy level to a lower one, the energy is emitted as a

#### Energy of emitted photon = $h\mathbf{n} = E_i - E_f$

- From Postulate 1,

photon.

$$\mathbf{E}_{i} = -\frac{\mathbf{R}_{h}}{\mathbf{n}_{i}^{2}} \qquad \mathbf{E}_{f} = -$$

# The Bohr Theory of the Hydrogen Atom

#### Bohr's Postulates

– If we make a substitution into the previous equation that states the energy of the emitted photon,  $h\,\nu$ , equals  $E_i$ - $E_f,$ 

$$\mathbf{h}\mathbf{n} = \mathbf{E}_{i} - \mathbf{E}_{f} = \left(-\frac{\mathbf{R}_{h}}{\mathbf{n}_{e}^{2}}\right) - \left(-\frac{\mathbf{R}_{h}}{\mathbf{n}_{e}^{2}}\right)$$

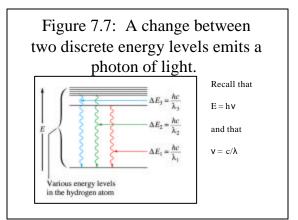
Rearranging, we obtain

$$h\mathbf{n} = \mathbf{R}_{h} \Big( \frac{1}{n_{f}^{2}} - \frac{1}{n_{i}^{2}} \Big)$$

#### A Problem to Consider

• Calculate the energy of a photon of light emitted from a hydrogen atom when an electron falls from level n = 3 to level n = 1.

$$E = hn = R_h \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$
$$E = (2.18 \cdot 10^{-18} \text{J}) \left(\frac{1}{1^2} - \frac{1}{3^2}\right)$$
$$E = 1.94 \cdot 10^{-18} \text{J}$$



Energy Changes in the Hydrogen Atom  $\Delta E = E_{\text{final state}} - E_{\text{initial state}}$  $I = \frac{hc}{DE}$ 

# The Bohr Theory of the Hydrogen Atom

· Bohr's Postulates

- Bohr's theory explains not only the emission of light, but also the absorption of light.
- When an electron falls from n = 3 to n = 2 energy level, a photon of red light (wavelength, 685 nm) is emitted.
- When red light of this same wavelength shines on a hydrogen atom in the n = 2 level, the energy is gained by the electron that undergoes a transition to n = 3.

"A curious thing about the atom model work of Bohr, prior to 1923 or 1924, was that if you look at the then-current papers you get the impression that everybody in the world was terrifically excited about the Bohr model and believed in it hook, line, and sinker, including the electron orbits as they are used in the ads for the atomic age nowadays. Bohr, on the other hand, was constantly making remarks, speeches, and admonitions to the effect that this is temporary and we ought to be looking for a way to do it right."

> - Edward U. Condon, to the Philosophical Society of Washington, December 2, 1960.

# **Quantum Mechanics**

- Bohr's theory established the concept of atomic energy levels but did not thoroughly explain the "wave-like" behavior of the electron.
- Current ideas about atomic structure depend on the principles of **quantum mechanics**, a theory that applies to subatomic particles such as electrons.

# **Quantum Mechanics**

- Bohr's theory established the concept of atomic energy levels but did not thoroughly explain the "wave-like" behavior of the electron.
- Current ideas about atomic structure depend on the principles of **quantum mechanics**, a theory that applies to subatomic particles such as electrons.

#### **Quantum Mechanics**

- The first clue in the development of quantum theory came with the discovery of the **de Broglie** relation
- relation. – In 1923, Louis de Broglie reasoned that if light exhibits particle aspects, perhaps particles of matter show characteristics of waves.
- He postulated that a particle with mass m and a velocity v has an associated wavelength.
- The equation 1 = h/mv is called the de Broglie relation.

# **Quantum Mechanics**

- If matter has wave properties, why are they not commonly observed?
- The de Broglie relation shows that a baseball (0.145 kg) moving at about 60 mph (27 m/s) has a wavelength of about 1.7 x 10<sup>-34</sup> m.

# $\mathbf{l} = \frac{6.63^{\circ}10^{\cdot34}\frac{\text{kgm}^2}{\text{s}}}{(0.145\text{kg})(27\text{m/s})} = 1.7^{\circ}10^{-34}\text{m}$

 This value is so incredibly small that such waves cannot be detected.

# **Quantum Mechanics**

- If matter has wave properties, why are they not commonly observed? For most objects, the wavelength is outside emr spectrum!
- Electrons have wavelengths on the order of a few picometers (1 pm =  $10^{-12}$  m). X-ray region of emr spectrum
- Under the proper circumstances, the wave character of electrons should be observable.

#### The Wavelength of an Electron

What is the wavelength of an electron moving at 6 x 10<sup>6</sup> m/s?

Using deBroglie's equation:

 $\begin{array}{l} \lambda = h/mv = (6.63 \ x \ 10^{-34} J\text{-s})/[(9.11 x \ 10^{-31} kg)(6 \ x \ 10^6 \ m/s)] \\ = 1.21 \ x \ 10^{-10} \ m \ or \ 0.121 \ nm \end{array}$ 

Remember that  $1 J = 1 \text{ kg-} \text{m}^2/\text{s}^2$ 

What is the energy associated with this electron?

 $E=hv=hc/\lambda=(6.63\ x\ 10^{-34}\ J\text{-s})(3.0\ x\ 10^8\ m/s)/(1.21\ x\ 10^{-10}\ m\ )$   $=\ 1.64\ x\ 10^{-15}\ J$ 

# **Quantum Mechanics**

• Heisenberg's uncertainty principle is a relation that states that the product of the uncertainty in position  $(\Delta x)$  and the uncertainty in momentum  $(m\Delta v_x)$  of a particle is a minimum of **h/4p.** 



 When *m* is large (for example, a baseball) the uncertainties are small, but for electrons, high uncertainties make it impossible to predict an exact orbit.

# Heisenberg Uncertainty Principle

 $\Delta x \cdot \Delta(mv) \geq \frac{h}{4p}$ 

- x = position
- mv = momentum
- h =Planck's constant
- The more accurately we know a particle's position, the less accurately we can know its momentum.

#### The Meaning of the Heisenberg Uncertainty Principle

We know that the electron has a mass of  $9.11 \times 10^{-31}$ kg and moves with an average speed of  $5 \times 10^6$  m/s in the H atom. The velocity is the only uncertainty in the momentum so we can calculate the uncertainty in the electrons position in the H atom using Heisenberg's Principle:

 $\begin{array}{l} \Delta x = h/(4\pi m \Delta v) = (6.63 x 10^{-34} J\text{-s}) / [4\pi (9.11 \; x \; 10^{-31} kg \;) (5 \; x 10^6 \; m/s)] \\ = \; 1 \; x \; 10^{-9} \; m \; (or \; 1 \; nm) \end{array}$ 

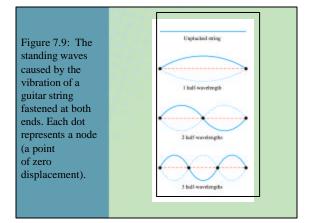
The diameter of an H atom is about 0.2 nm, so the uncertainty in the electron's position is greater than the size of the H atom! We simply can't say where the electron is "located" in the H atom. Now, try the calculation using a 70.0 kg man running at 10.0 m/s! You will find that the uncertainty in the man's position is Very small.

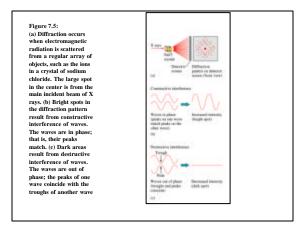
# Quantum Mechanics

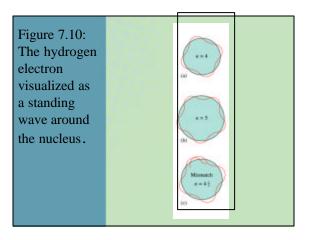
Based on the wave properties of the atom

#### $\hat{H}\psi = E\psi$

- $\Psi$  = wave function
- $\hat{H}$  = mathematical operator
- E = total energy of the atom
- A specific wave function is often called an orbital.







# **Quantum Mechanics**

- Although we cannot precisely define an electron's orbit, we can obtain the *probability* of finding an electron at a given point around the nucleus.
  - Erwin Schrodinger defined this probability in a mathematical expression called a wave function, denoted y (psi).
  - The probability of finding a particle in a region of space is defined by y<sup>2</sup>. (see Figures 7.18 and 7.19)

### Heisenberg's Uncertainty Principle:

 $(? x)(? p_x) = h/4p$  or  $(? x)(? v_x) = h/4pm$ ,  $h = 6.63 x 10^{-34} J_s$ For subatomic particles which have very small masses, the uncertainties in position and speed can be quite large. We can only talk about the probability, ?<sup>2</sup>, of where we might find them.

# **Probability Distribution**

- square of the wave function
- probability of finding an electron at a given position
- Radial probability distribution is the probability distribution in each spherical shell.

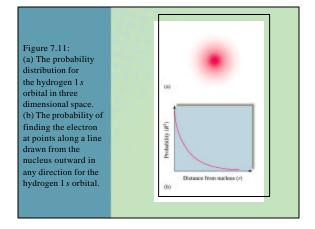
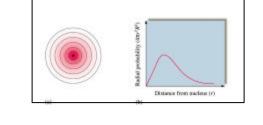
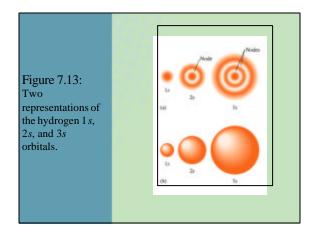
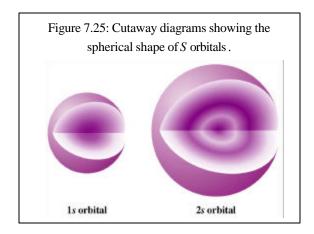


Figure 7.12: (a) Cross section of the hydrogen 1s orbital probability distribution divided into successive thin spherical shells. (b) The radial probability distribution.

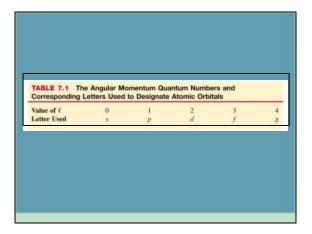






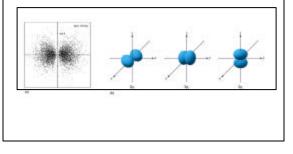
# Quantum Numbers (QN)

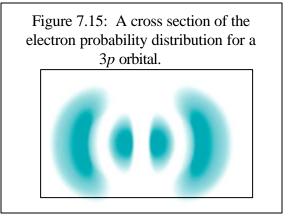
- 1. Principal QN (*n* = 1, 2, 3, . . .) related to size and energy of the orbital.
- 2. Angular Momentum QN (l=0 to n-1) relates to shape of the orbital.
- 3. Magnetic QN (m<sub>1</sub> = 1 to -l) relates to orientation of the orbital in space relative to other orbitals.
- 4. Electron Spin QN  $(m_s = +\frac{1}{2}, -\frac{1}{2})$  relates to the spin states of the electrons.

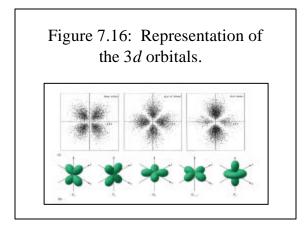


п	ε	Orbital Designation	m,	Number of Orbitab
1	0	Lr.	0	1
2	0	2/	0	1
	1	20	-1, 0, +1	3
3	0	31	0	1
	1	hp 3d	-1, 0, 1	3
	2	3d	-2, -1, 0, 1, 2	. 5
4	0	41	0	1.
		40	-1, 0, 1	3 5 T
	2	4.1	-2, -1, 0, 1, 2	5
	3	45	-3, -2, -1, 0, 1, 2, 3	7

Figure 7.14: Representation of the 2p orbitals . (a) The electron probability distributed for a 2p orbital. (b) The boundary surface representations of all three 2p orbitals .







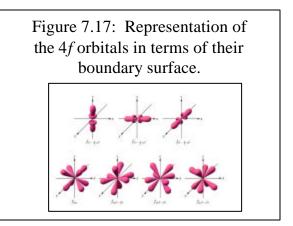


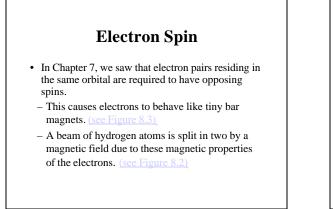
Figure 7.18: Orbital energy levels for the hydrogen atom.

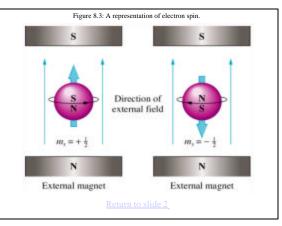
$$E \begin{vmatrix} 3s - 3p - - - 3d \\ 2s - 2p - - - \\ 1s - \end{vmatrix}$$

In the quantum (wave) mechanical model, the electron is a "standing wave" and can be described by a series of wave functions (orbitals) That describe the possible energies and spatial distributions available to the electron. From the Heisenberg Uncertainty Principle, the model cannot predict the exact electron motions – the  $\Psi^2$  represents the probability distribution of the electrons in that orbital. We depict the orbital as an electron density map. The size of an orbital is arbitrarily defined as the surface that contains 90% of the total electron density.

A Summary of the H Atom

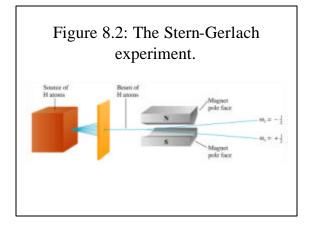
The H atom has many types of orbitals – in the ground state, the single electron resides in the 1s orbital. The electron can be excited to higher energy orbitals if the appropriate amount of energy is put into the atom.





#### I Electronic Structure of Atoms

- A Electron Spin and the Pauli Exclusion Principle
  - 1 The Stern-Gerlach Experiment demonstrated interaction between the magnetic field of the spinning electron (in a stream of Ag atoms) and an external magnetic field. When  $m_s =$ +  $\frac{1}{2}$  same poles are aligned; when  $m_s = -\frac{1}{2}$  opposite poles are aligned



# Pauli Exclusion Principle

- In a given atom, no two electrons can have the same set of four quantum numbers  $(n, l, m_l, m_s)$ .
- *Therefore*, an orbital can hold only two electrons, and they must have opposite spins.

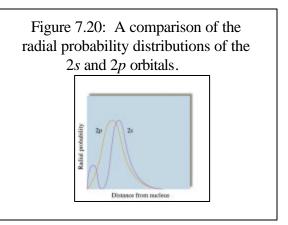
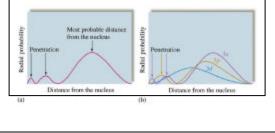
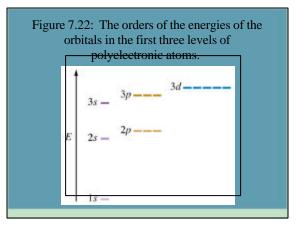
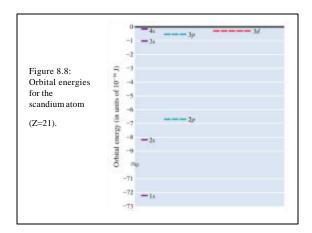
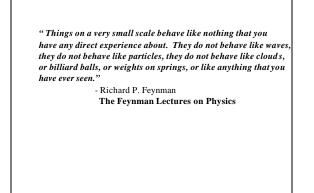


Figure 7.21: (a) The radial probability distribution for an electron in a 3*s* orbital. (b) The radial probability distribution for the 3*s*, 3*p*, and 3*d* orbitals.









# Aufbau Principle

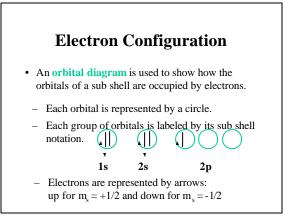
- The Aufbau principle is a scheme used to reproduce the ground state electron configurations of atoms by following the "building up" order.
- Listed below is the order in which all the possible subshells fill with electrons.
- 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f
  - You need not memorize this order. As you will see, it can be easily obtained.

### Aufbau Principle

• As protons are added one by one to the nucleus to build up the elements, electrons are similarly added to these hydrogen-like orbitals.

# **Electron Configuration**

- An "electron configuration" of an atom is a particular distribution of electrons among available sub shells.
- The notation for a configuration lists the subshell symbols sequentially with a superscript indicating the number of electrons occupying that sub shell.
- For example, lithium (atomic number 3) has two electrons in the "1s" sub shell and one electron in the "2s" sub shell 1s<sup>2</sup> 2s<sup>1</sup>.



# The Pauli Exclusion Principle

- The Pauli exclusion principle, which summarizes experimental observations, states that *no two electrons can have the same four quantum numbers*.
  - In other words, an orbital can hold <u>at most</u> two electrons, and then <u>only if the electrons have</u> <u>opposite spins</u>.

### The Pauli Exclusion Principle

• The maximum number of electrons and their orbital diagrams are:

		-	
Sub shell	Number of Orbitals	Maximum Number of Electrons	
s $(l = 0)$	1	2	(1)
p ( <i>l</i> = 1)	3	6	
d ( <i>l</i> =2)	5	10	
f $(l=3)$	7	14	

### Hund's Rule

• The lowest energy configuration for an atom is the one having the maximum number of unpaired electrons allowed by the Pauli principle in a particular set of degenerate orbitals.

#### **Orbital Diagrams**

- Consider carbon (Z = 6) with the ground state configuration  $1s^{2}2s^{2}2p^{2}$ .
  - Three possible arrangements are given in the following orbital diagrams.

<u>(1)</u>

 $\square$ 

<u>()</u>

Diagram 1:

Diagram 2:

Diagram 3:

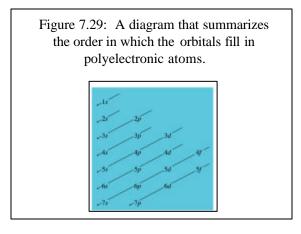
 Each state has a different energy and different magnetic characteristics.

 $\hat{\Pi}$ 

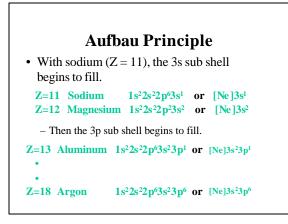
# Orbital Diagrams Hund's rule states that the lowest energy arrangement (the "ground state") of electrons in a sub-shell is obtained by putting electrons into <u>separate orbitals</u> of the sub shell with the same spin before pairing electrons. Looking at carbon again, we see that the ground state configuration corresponds to diagram 1 when following Hund's rule.

# **Magnetic Properties**

- Although an electron behaves like a tiny magnet, two electrons that are opposite in spin cancel each other. Only atoms with unpaired electrons exhibit magnetic susceptibility.
  - A paramagnetic substance is one that is weakly attracted by a magnetic field, usually the result of <u>unpaired electrons</u>.
  - A diamagnetic substance is not attracted by a magnetic field generally because it has only paired electrons.

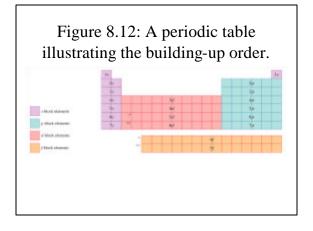


	Aufba	au Prin	cip	le
• With bo filling th	ron (Z=5) ne 2p subs		rons	begin
Z=5	Boron	$1s^22s^22p^1$	or	[He]2s <sup>2</sup> 2p <sup>1</sup>
Z=6	Carbon	$1s^22s^22p^2$	or	[He]2s <sup>2</sup> 2p <sup>2</sup>
<b>Z=7</b>	Nitrogen	$1s^22s^22p^3$	or	[He]2s <sup>2</sup> 2p <sup>3</sup>
Z=8	Oxygen	$1s^22s^22p^4$	or	[He]2s <sup>2</sup> 2p <sup>4</sup>
Z=9	Fluorine	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>5</sup>	or	[He]2s <sup>2</sup> 2p <sup>5</sup>
7. 10	Noon	1 02 20 22 20	07	[He]2s <sup>6</sup> 2p <sup>6</sup>



# Configurations and the Periodic Table

- Electrons that *reside in the outermost shell* of an atom or in other words, those electrons outside the "noble gas core" are called valence electrons.
  - These electrons are primarily involved in chemical reactions.
  - Elements within a given group have the same "valence shell configuration."
  - This accounts for the similarity of the chemical properties among groups of elements.



# Valence ElectronsThe electrons in the outermost principlequantum level of an atom.AtomValence ElectronsCa2N5Br7

Inner electrons are called core electrons.

# Configurations and the Periodic Table

- The following slide illustrates how the periodic table provides a sound way to remember the Aufbau sequence.
  - In many cases you need only the configuration of the outer elements.
  - You can determine this from their position on the periodic table.
  - The total number of valence electrons for an atom equals its group number.

Figure 7.26: Electron configurations for potassium through krypton.

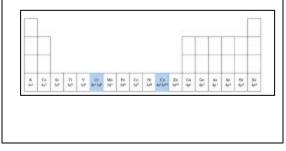
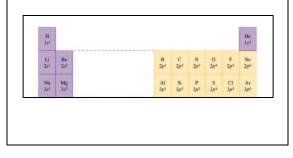


Figure 7.25: The electron configurations in the type of orbital occupied last for the first 18 elements.



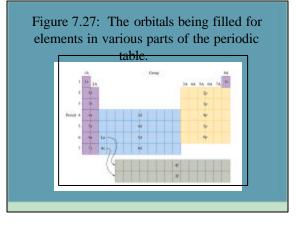
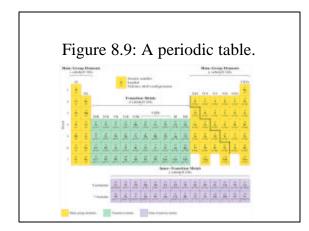


Figure 7.24: Mendeleev's early periodic table, published in 1872. Note the spaces left for missing elements with atomic masses 44, 68, 72, 100.

				TABEL	LE IL			
*0+0*	810	RD COLOR	RIG?	201010 10 2011	Ring a	(18)(16)(14) 第162 第163	RM RM RHDT	oniore an
1		and a				-	1.4	
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	Town and	8		78	49 - 19	60.00	81-00	100000000000000000000000000000000000000
٠	84-85	80-40	1-1-10	2++85	****	*1-188		Pare U.A. Dorothe.
۰.	144-108	24-114	30,4168	80004	84+123	101125	0+423	
	84-188	88 + ISF	101-108	718-148				
٠	- Fel		-	1.1.1.1.1.1			-	and the second
-	-	-	7Er+178	714-185	TATINE	*****	-	101 - 405, 73 - 187, 18 - 195, 75 - 187,
-	14-10-14-16	141-2.00	11-25-	44-223	8- + 254	-		
		-	-		-	01.040 V	-	

Mendeleev and as Actu	on of the Properties of German ally Observed	num as Predicted by
Properties of Germanium	Predicted In 1871	Observed in 1886
Atomic weight	72	72.3
Density	5.5 g/cm <sup>3</sup>	5.47 g/cm <sup>3</sup>
Specific heat	0.31 B(°C+2)	0.32 JA*C · g)
Melting point	Very high	960°C
Oxide formula	RO <sub>2</sub>	GeO1
Oxide density	4.7 g/cm <sup>1</sup>	4.70 g/cm <sup>2</sup>
Chloride formula	RCL <sub>4</sub>	GeCla
bp of chloride	100°C	86%C



		Element 113	Property
Lead	1	Thallium	Chemically like
298		297	Atomic mass
14 g/ml.	3	16 g/ml.	Density
70°C		430°C	Melting point
150°C		1100°C	
		16 g/ml. 430°C 1100°C	Density Melting point Boiling point

# Information Contained in the Periodic Table

- 1. Each group member has the same valence electron configuration (these electrons primarily determine an atom's chemistry).
- 2. The electron configuration of any representative element.
- 3. Certain groups have special names (alkali metals, halogens, etc).
- 4. Metals and nonmetals are characterized by their chemical and physical properties.

# Configurations and the Periodic Table

- Note that elements within a given family have similar configurations.
  - The Group IIA elements are sometimes called the *alkaline earth metals*.

 Beryllium
 1 s<sup>2</sup>2s<sup>2</sup>

 Magnesium
 1 s<sup>2</sup>2 s<sup>2</sup>2 p<sup>6</sup>3s<sup>2</sup>

 Calcium
 1 s<sup>2</sup>2 s<sup>2</sup>2 p<sup>6</sup>3 s<sup>2</sup>3 p<sup>6</sup>4s<sup>2</sup>

# Configurations and the Periodic Table

• Note that elements within a given family have similar configurations.

- For instance, look at the *noble gases*.

 Helium
 1s<sup>2</sup>

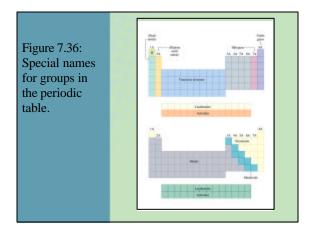
 Neon
 1s<sup>2</sup>2s<sup>2</sup>2p<sup>6</sup>

 Argon
 1s<sup>2</sup>2s<sup>2</sup>2p<sup>6</sup>3s<sup>2</sup>3p<sup>6</sup>

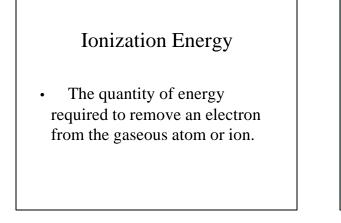
 Krypton
 1s<sup>2</sup>2s<sup>2</sup>2p<sup>6</sup>3s<sup>2</sup>3p<sup>6</sup>3d<sup>10</sup>4s<sup>2</sup>4p<sup>6</sup>

# Broad Periodic Table Classifications

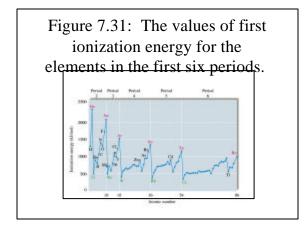
- Representative Elements (main group): filling *s* and *p* orbitals (Na, Al, Ne, O)
- Transition Elements : filling *d* orbitals (Fe, Co, Ni)
- Lanthanide and Actinide Series (inner transition elements): filling 4*f* and 5*f* orbitals (Eu, Am, Es)

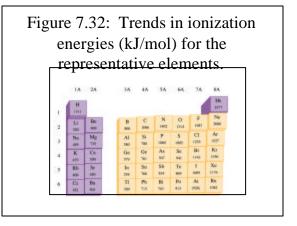


Dever	Valence Bestiran Config- antitice	Density at 25%	29	10	First Kongution Esergy (ku/wst)	Atomic (cavalent) Factus (pre)	Radius Radius Gard
L4 Na	30 <sup>2</sup> 30 <sup>2</sup>	8.59 8.97	140- 140	(330) 890	1030 415	157	91 93
14.20	44' 54' 54'	888 1.19 1.87	54 39 29	700 868 890	419 60 62	207	130



į,	Elements							1993.00
5	Bernent	1	4	12	1.	4	4	Ir:
	Na	495	4550					
1	Mg	735	1445	7730	Core el	octrons*		
	AL	580	1815	2740	11,600			
	SI	780	1575	3220	4350	16.100		
	P	1060	1890	2905	4950	6230	21,200	
	5	1005	2260	3375	4565	6950	8490	27,000
	CI	1255	2295	3850	5160	6560	9360	11,000
	Ar	1527	3965	3945	5770	7230	8780	12,000
	"Note the large core electrons	r jump in is	elisation cei		g from rame d increase -	ral of valence	dictrons to n	onesal of





# Periodic Trends

- First ionization energy:
- increases from left to right across a period;
- decreases going down a group.

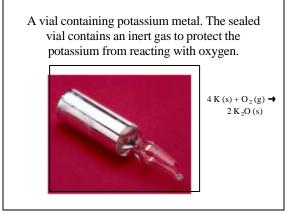
This means that metallic character increases as we go from top to bottom of a group

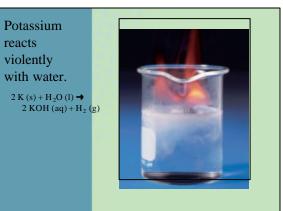
	First Ionization he Alkali Metals ses	
Atom	I, (kJ/mol)	
Group 1A		
Li	520	
Na	495	
K	419	
Rb	409	
Cs	382	
Group 8A		
He	2377	
Ne	2068	
Ar	1527	
Kr	1356	
Xa	1176	
Rn	1042	

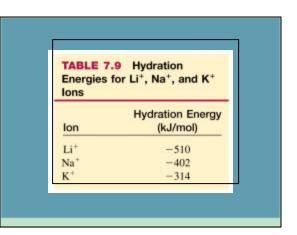
Sodium metal is so reactive that it is stored under kerosene to protect it from the oxygen in the air.

 $4 \operatorname{Na}(s) + \operatorname{O}_{2}(g) \rightarrow 2 \operatorname{Na}_{2}\operatorname{O}(s)$ 









# **Electron Affinity**

• The energy change associated with the addition of an electron to a gaseous atom.

•  $X(g) + e^- \rightarrow X^-(g)$ 

# **Periodic Properties**

- Electron Affinity
  - The electron affinity is the energy change for the process of adding an electron to a neutral atom in the gaseous state to form a negative ion.

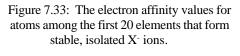
-For a chlorine atom, the first electron affinity is illustrated by:

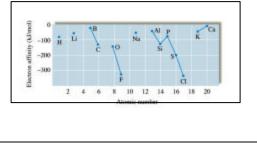
 $\begin{array}{c} Cl([Ne]3s^23p^5) + e^- @ Cl^-([Ne]3s^23p^6) \\ & Electron Affinity = -349 \text{ kJ/mol} \end{array}$ 

## **Periodic Properties**

- Electron Affinity
- The more negative the electron affinity, the more stable the negative ion that is formed.
- Broadly speaking, the general trend goes from lower left to upper right as electron affinities become more negative.
- Table 8.4 gives the electron affinities of the main-group elements.

Period	IA	IIIA	IVA	VA	VIA	VIIA
1	н					
	-73					
2	ы	В	с	N	0	F
	-60	-27	-122	0	-141	-328
3	Na	AI	Si	Р	8	CI
	-53	-44	-134	-72	-200	-349
4	K	Ga	Ge	As	Se	Br
	-48	-30	-120	-77	-195	-325
5	Rb	In	Sn	Sb	Te	1
	-47	-30	-121	-101	-190	-295
6	Cs	п	Pb	Bi	Po	At
	-45	-30	-110	-110	-180	-270





A CONTRACTOR OF A CONTRACTOR O	7 Electron of the Halogens
Atom	Electron Affinity (kJ/mol)
F	-327.8
CI	-348.7
Br	-324.5
I	-295.2

# Periodic Trends

- Atomic Radii:
- decrease going from left to right across a period;
- increase going down a group.

