

Nerve Signal Transmission	
	3
	Li
<ul> <li>Movement of ions across cell memoranes</li> </ul>	6.94
is the basis for the transmission of nerve	11
signals.	Na
<ul> <li>Na<sup>+</sup> and K<sup>+</sup> ions are numbed across</li> </ul>	22.99
mambrance in especite directions through	19
	K
ion channels.	39.10
<ul> <li>– Na<sup>+</sup> out and K<sup>+</sup> in</li> </ul>	37
<ul> <li>The ion channels can differentiate Na<sup>+</sup></li> </ul>	Rb
from $K^+$ by their difference in size	85.47
for the by their difference in size.	55
<ul> <li>Ion size and other properties of atoms are</li> </ul>	Cs
periodic properties—properties whose	132.91
values can be predicted based on the	87
element's position on the periodic table	Fr
ciement a position on the periodic table.	(223.02)

### Mendeleev (1834-1907)

- · Ordered elements by atomic mass
- · Saw a repeating pattern of properties
- Periodic law: When the elements are arranged in order of increasing atomic mass, certain sets of properties recur periodically.



- Put elements with similar properties in the same column
- Used pattern to predict properties of undiscovered elements
- Where atomic mass order did not fit other properties, he reordered by other properties.

Te and I





- Mendeleev's periodic law allows us to predict *what* the properties of an element will be based on its position on the table.
- It doesn't explain why the pattern exists.
- Quantum mechanics is a theory that explains why the periodic trends in the properties exist.
   – Knowing why allows us to predict what.





### **Electron Spin**

- Experiments by Stern and Gerlach showed that a beam of silver atoms is split in two by a magnetic field.
- The experiment reveals that the electrons spin on their axis.
- As they spin, they generate a magnetic field.
   Spinning charged particles generates a magnetic field.
- If there is an even number of electrons, about half the atoms will have a net magnetic field pointing "north" and the other half will have a net magnetic field pointing "south."



# Spin Quantum Number, *m<sub>s</sub>*, and Orbital Diagrams

- $m_s$  can have values of  $+\frac{1}{2}$  or  $-\frac{1}{2}$ .
- Orbital diagrams use a square to represent each orbital and a half-arrow to represent each electron in the orbital.
- By convention, a half-arrow pointing up is used to represent an electron in an orbital with spin up.
- Spins must cancel in an orbital.
  - Paired





### **Pauli Exclusion Principle**

- Knowing the number of orbitals in a sublevel allows us to determine the maximum number of electrons in the sublevel:
  - s sublevel has 1 orbital; therefore, it can hold 2 electrons.
  - *p* sublevel has 3 orbitals; therefore, it can hold
    6 electrons.
  - *d* sublevel has 5 orbitals; therefore, it can hold 10 electrons.
  - *f* sublevel has 7 orbitals; therefore, it can hold 14 electrons.





# Sublevel Splitting in Multielectron Atoms The sublevels in each principal energy shell of hydrogen all have the same energy or other single electron systems. We call orbitals with the same energy degenerate. For multielectron atoms, the energies of the sublevels are split. Caused by charge interaction, shielding, and penetration The lower the value of the *l* quantum number, the less energy the sublevel has. *s* (*l* = 0) < *p* (*l* = 1) < *d* (*l* = 2) < *f* (*l* = 3)

### **Coulomb's Law**

$$E = \frac{1}{4\pi\varepsilon_0} \frac{q_1 q_2}{r}$$

- Coulomb's law describes the attractions and repulsions between charged particles.
- For like charges, the potential energy (*E*) is positive and decreases as the particles get farther apart as *r* increases.
- For opposite charges, the potential energy is negative and becomes more negative as the particles get closer together.
- The strength of the interaction increases as the size of the charges increases.
  - Electrons are more strongly attracted to a nucleus with a 2+ charge than a nucleus with a 1+ charge.

© 2017 Pearson Education, Inc.

### **Shielding and Effective Nuclear Charge**

- Each electron in a multielectron atom experiences both the attraction to the nucleus and the repulsion by other electrons in the atom.
- These repulsions cause the electron to have a net reduced attraction to the nucleus; it is shielded from the nucleus.
- The total amount of attraction that an electron feels for the nucleus is called the **effective nuclear charge** of the electron.







# Effect of Penetration and Shielding Penetration causes the energies of sublevels in the same principal level to not be degenerate. In the fourth and fifth principal levels, the effects of penetration become so important that the *s* orbital ies lower in energy than the *d* orbitals of the previous principal level. The energy separations between one set of orbitals and the next become smaller beyond the 4s. The ordering can therefore vary among elements, causing variations in the electron configurations of the transition metals and their ions.



### **Filling the Orbitals with Electrons**

- Energy levels and sublevels fill from lowest energy to highest:
  - $\quad s \to p \to d \to f$
  - Aufbau principle
- Orbitals that are in the same sublevel have the same energy.
- No more than two electrons per orbital
  - Pauli exclusion principle
- When filling orbitals that have the same energy, place one electron in each before completing pairs.
  - Hund's rule

# Electron Configuration of Atoms in Their Ground State

• The electron configuration is a listing of the sublevels in order of filling with the number of electrons in that sublevel written as a superscript.

Kr = 36 electrons =  $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$ 

 A short-hand way of writing an electron configuration is to use the symbol of the previous noble gas in brackets [] to represent all the inner electrons and then just write the last set.

Rb = 37 electrons =  $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1 = [Kr]5s^1$ 





1A 1 H	Duter Ele	ectron Co	onfigura	ations o	f Eleme	nts 1–1	8 8A 2 He
$1s^1$	2A	3A	4A	5A	6A	7A	1 <i>s</i> <sup>2</sup>
$3$ <b>Li</b> $2s^1$	$\begin{array}{c} 4\\ \mathbf{Be}\\ 2s^2 \end{array}$	$ \begin{array}{c} 5\\ \mathbf{B}\\ 2s^22p^1 \end{array} $	$6$ C $2s^2 2p^2$	$7$ <b>N</b> $2s^2 2p^3$	8 0 $2s^22p^4$	9 F 2s <sup>2</sup> 2p <sup>5</sup>	10 Ne $2s^22p^6$
11 <b>Na</b> 3s <sup>1</sup>	$ \begin{array}{c} 12 \\ \mathbf{Mg} \\ 3s^2 \end{array} $	$ \begin{array}{c} 13\\ \textbf{Al}\\ 3s^23p^1 \end{array} $	$ \begin{array}{r} 14 \\ \mathbf{Si} \\ 3s^2 3p^2 \end{array} $	$ \begin{array}{c} 15 \\ \mathbf{P} \\ 3s^2 3p^3 \end{array} $	$ \begin{array}{r} 16\\ \mathbf{S}\\ 3s^23p^4 \end{array} $	$ \begin{array}{c} 17\\ \mathbf{Cl}\\ 3s^23p^5 \end{array} $	$     18 \\     Ar \\     3s^2 3p^6 $



- The group number corresponds to the number of valence electrons.
- The length of each "block" is the maximum number of electrons the sublevel can hold.
- The period number corresponds to the principal energy level of the valence electrons.





### **Irregular Electron Configurations**

- We know that, because of sublevel splitting, the 4*s* sublevel is lower in energy than the 3*d*; therefore, the 4*s* fills before the 3*d*.
- But the difference in energy is not large.
- Some of the transition metals have irregular electron configurations in which the *ns* only partially fills before the (*n*−1)*d* or doesn't fill at all.
- Therefore, their electron configuration must be found experimentally.

## Irregular Electron Configurations

- Expected
- Cr =  $[Ar]4s^23d^4$
- Cu =  $[Ar]4s^23d^9$
- Mo =  $[Kr]5s^24d^4$
- Ru = [Kr] $5s^24d^6$
- Pd = [Kr] $5s^24d^8$

- Found experimentally
- Cr =  $[Ar]4s^{1}3d^{5}$
- Cu =  $[Ar]4s^{1}3d^{10}$
- Mo =  $[Kr]5s^{1}4d^{5}$
- Ru = [Kr] $5s^{1}4d^{7}$
- $Pd = [Kr]5s^{0}4d^{10}$















# Electron Configuration of Anions in Their Ground State

## Anions are formed when nonmetal atoms gain enough electrons to have eight valence electrons. – Filling the *s* and *p* sublevels of the valence shell

- The sulfur atom has six valence electrons.
   S atom = 1s<sup>2</sup>2s<sup>2</sup>2p<sup>6</sup>3s<sup>2</sup>3p<sup>4</sup>
- To have eight valence electrons, sulfur must gain two more.

 $S^{2-}$  anion =  $1s^22s^22p^63s^23p^6$ 





















# Quantum-Mechanical Explanation for the Period Trend in Atomic Radius

- The larger the effective nuclear charge an electron experiences, the stronger the attraction it will have for the nucleus.
- The stronger the attraction the valence electrons have for the nucleus, the closer their average distance will be to the nucleus.
- Traversing across a period increases the effective nuclear charge on the valence electrons.
- Quantum-mechanics predicts the atoms should get smaller across a period.





### **Electron Configurations of Transition Metal Cations in Their Ground State** When transition metals form cations, the first electrons removed are the valence electrons, even though other electrons were added after. • Electrons may also be removed from the sublevel closest to the valence shell after the valence electrons. The iron atom has two valence electrons: Fe atom = $1s^22s^22p^63s^23p^64s^23d^6$ When iron forms a cation, it first loses its valence electrons: Fe<sup>2+</sup> cation = $1s^22s^22p^63s^23p^63d^6_{\text{Fe}^{3+}}$ [Ar] • It can then lose 3d electrons: $Fe^{3+}$ cation = $1s^22s^22p^63s^23p^63d^5$ © 2017 Pearson Education. Inc.

### Magnetic Properties of Transition Metal Atoms and Ions

- Electron configurations that result in unpaired electrons mean that the atom or ion will have a net magnetic field; this is called **paramagnetism**.
   Will be attracted to a magnetic field
- Electron configurations that result in all paired electrons mean that the atom or ion will have no magnetic field; this is called **diamagnetism**.
  - Slightly repelled by a magnetic field

© 2017 Pearson Education, Inc.

### **Trends in Ionic Radius**

- lons in the same group have the same charge.
- Ion size increases down the column.
  - Higher valence shell, larger ion
- Cations are smaller than neutral atoms; anions are larger than neutral atoms.
- Cations are smaller than anions.
  - Except Rb<sup>+</sup> and Cs<sup>+</sup>, bigger or same size as F<sup>-</sup> and O<sup>2-</sup>
- Larger positive charge = smaller cation
  - For isoelectronic species
  - Isoelectronic = same electron configuration
- Larger negative charge = larger anion
  - For isoelectronic species



# Explanation for the Trends in Cation Radius When atoms form cations, the valence electrons are removed. The farthest electrons from the nucleus are the *p* or *d* electrons in the (*n* – 1) energy level. This results in the cation being smaller than the atom.

















# Explanation for the Trends in First Ionization Energy

- The strength of attraction is related to the most probable distance the valence electrons are from the nucleus and the effective nuclear charge the valence electrons experience.
- The larger the orbital an electron is in, the farther its most probable distance will be from the nucleus and the less attraction it will have for the nucleus.
- Quantum-mechanics predicts the atom's first ionization energy should get lower down a column.

© 2017 Pearson Education, Inc.

# Explanation for the Trends in First Ionization Energy

- Traversing across a period increases the effective nuclear charge on the valence electrons.
- Quantum-mechanics predicts the atom's first ionization energy should get larger across a period.









### **Electron Affinity**

- Energy is released when a neutral atom gains an electron.
  - Gas state
  - M(g) + 1e<sup>-</sup>  $\rightarrow$  M<sup>1-</sup>(g) + EA
- Electron affinity is defined as exothermic (-) but may actually be endothermic (+).
  - Some alkali earth metals and all noble gases are endothermic. Why?
- The more energy that is released, the larger the electron affinity.
  - The more negative the number, the larger the EA.

© 2017 Pearson Education, Inc.

# **Trends in Electron Affinity**Alkali metals decrease electron affinity down the column. But not all groups do Generally irregular increase in EA from second period to third period "Generally" increases across period Becomes more negative from left to right Not absolute Group 5A generally lower EA than expected because extra electron must pair Groups 2A and 8A generally very low EA because added electron goes into higher energy level or sublevel Highest EA in any period = halogen

1A							8A
Н -73	2A	3A	4A	5A	6A	7A	<b>He</b> >0
Li	<b>Be</b> >0	<b>B</b>	<b>C</b>	N	<b>O</b>	<b>F</b>	<b>Ne</b>
-60		-27	-122	>0	-141	-328	>0
<b>Na</b>	<b>Mg</b>	Al	<b>Si</b>	Р	<b>S</b>	Cl	<b>Ar</b>
-53	>0	-43	-134	-72	-200	-349	>0
К	<b>Ca</b>	<b>Ga</b>	<b>Ge</b>	<b>As</b>	<b>Se</b>	<b>Br</b>	<b>Kr</b> >0
-48	-2	-30	-119	-78	-195	-325	
<b>Rb</b>	<b>Sr</b>	<b>In</b>	<b>Sn</b>	<b>Sb</b>	<b>Te</b>	I	<b>Xe</b> >0
-47	-5	-30	-107	-103	-190	-295	

## **Properties of Metals and Nonmetals**

- Metals
  - Malleable and ductile
  - Shiny, lustrous, reflect light
  - Conduct heat and electricity
  - Most oxides basic and ionic
  - Form cations in solution
  - Lose electrons in reactions-oxidized
- Nonmetals
  - Brittle in solid state
  - Dull, nonreflective solid surface
  - Electrical and thermal insulators
  - Most oxides are acidic and molecular
  - Form anions and polyatomic anions
  - Gain electrons in reactions—reduced



							1	Meta	llic c	hara	cter d	lecre	ases					_	
																		$\neg$	
		1A 1						11 -	Г			.1.							8A 18
	1	1	2A		Me			etano			onmet	ais		3A	4A	5A	6A	7A	2
	1	Η	2	1										13	14	15	16	17	He
ses	2	3 Li	4 Be											5 B	6 C	7 N	8	9 F	10 Ne
rea	2	11	12	3B	4B	5B	6B	7B		-8B-		1B	2B	13	14	15	16	17	18
Inc	ds.	Na	Mg	3	4	5	6	7	8	9	10	11	12	Al	Si	Р	S	Cl	Ar
er	oin 4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
raci	Pe	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
na	5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	Ι	Xe
2 2	6	55 Cs	56 Ba	57 1a	72 Hf	73 Ta	74 W	75 Re	76	77	78 Pt	79 <b>A</b> 11	80 Hg	81 T1	82 Ph	83 Bi	84 Po	85 At	86 Rn
Eall	7	87	88	89	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118
Me		Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cm		Fl		Lv		
			La	nthan	ides	58	59 D	60	61 D	62	63 F	64	65	66 D	67	68 F	69	70	71
					H	Ce	01	Na	Pm 02	<b>Sm</b>	Eu 05	Ga	1D 07	Dy	H0	Er	1m 101	102	Lu 102
				Actin	ides	Th	Pa	92 U	Np	94 Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

# Explanation for the Trends in Metallic Character

- Metals generally have smaller first ionization energies, and nonmetals generally have larger electron affinities.
  - Except for the noble gases
- ∴ quantum mechanics predicts the atom's metallic character should increase down a column because the valence electrons are not held as strongly.
- .: quantum mechanics predicts the atom's metallic character should decrease across a period because the valence electrons are held more strongly and the electron affinity increases.

© 2017 Pearson Education, Inc.

### **Trends in the Alkali Metals** Atomic radius increases down the column. Ionization energy decreases down the column. Very low ionization energies - Good reducing agents; easy to oxidize - Very reactive; not found uncombined in nature React with nonmetals to form salts - Compounds generally soluble in water : found in seawater Electron affinity decreases down the column. Melting point decreases down the column. - All very low MP for metals Density increases down the column. Except K - In general, the increase in mass is greater than the increase in volume. © 2017 Pearson Education Inc.

TABLE 8	<b>5.2</b> Properties of the Alka				
Element	Electron Configuration	Atomic Radius (pm)	IE <sub>1</sub> (kJ/mol)	Density at 25 °C (g/cm <sup>3</sup> )	Melting Point (°C)
Li	[He] 2s <sup>1</sup>	152	520	0.535	181
Na	[Ne] 3s <sup>1</sup>	186	496	0.968	102
к	[Ar] 4s <sup>1</sup>	227	419	0.856	98
Rb	[Kr] 5s <sup>1</sup>	248	403	1.532	39
Cs	[Xe] 6s <sup>1</sup>	265	376	1.879	29
	R	eactions of the All	kali Metals w	vith Water	1
	R	eactions of the All	kali Metals w	vith Water	
	R	eactions of the All	kali Metals w	vith Water	

## **Trends in the Halogens**

- Atomic radius increases down the column.
- Ionization energy decreases down the column.
- Very high electron affinities
   Good oxidizing agents; easy to reduce
  - Cood Unidizing agents, easy to reduce
  - Very reactive; not found uncombined in nature
  - React with metals to form salts
  - Compounds generally soluble in water ∴ found in seawater



TABLE 8.	3 Properties of the I	lalogens*				
Element	Electron Configuration	Atomic Radius (pm)	EA (kJ/mol)	Melting Point (°C)	Boiling Point (°C)	Density of Liquid (g/cm <sup>3</sup> )
F	[He] 2s <sup>2</sup> 2p <sup>5</sup>	72	-328	-219	-188	1.51
CI	[Ne] 3s <sup>2</sup> 3p <sup>5</sup>	99	-349	-101	-34	2.03
Br	[Ar] 4s <sup>2</sup> 4p <sup>5</sup>	114	-325	-7	59	3.19
I.	[Kr] 5s <sup>2</sup> 5p <sup>5</sup>	133	-295	114	184	3.96
		Chlor	rine			

### **Reactions of Alkali Metals with Halogens**

- Alkali metals are oxidized to the 1+ ion.
- Halogens are reduced to the 1- ion.
- The ions then attach together by ionic bonds.
- The reaction is exothermic.









TABLE 8	.4 Properties of the Nobl	e Gases*			
Element	Electron Configuration	Atomic Radius (pm)**	IE <sub>1</sub> (kJ/mol)	Boiling Point (K)	Density of Gas (g/L at STP
	<u>^</u>	30	2272	4.2	0.48
He	1s <sup>2</sup>	32	2312	4.2	0.18
He Ne	1s <sup>2</sup> [He] 2s <sup>2</sup> 2p <sup>6</sup>	70	2081	27.1	0.18
He Ne Ar	1s <sup>2</sup> [He] 2s <sup>2</sup> 2p <sup>6</sup> [Ne] 3s <sup>2</sup> 3p <sup>6</sup>	70 98	2081 1521	27.1 87.3	0.90
He Ne Ar Kr	1s <sup>2</sup> [He] 2s <sup>2</sup> 2p <sup>6</sup> [Ne] 3s <sup>2</sup> 3p <sup>6</sup> [Ar] 4s <sup>2</sup> 4p <sup>6</sup>	70 98 112	2081 1521 1351	27.1 87.3 119.9	0.18 0.90 1.78 3.74
He Ne Ar Kr Xe	1s <sup>2</sup> [He] 2s <sup>2</sup> 2p <sup>6</sup> [Ne] 3s <sup>2</sup> 3p <sup>6</sup> [Ar] 4s <sup>2</sup> 4p <sup>6</sup> [Kr] 5s <sup>2</sup> 5p <sup>6</sup>	70 98 112 130	2081 1521 1351 1170	27.1 87.3 119.9 165.1	0.18 0.90 1.78 3.74 5.86

© 2017 Pearson Education, Inc.

**Noble Gases**