



# General Chemistry

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# Topics Breakdown

## ● Lecture Plan

- Electromagnetic Radiation
- Photoelectric effect
- Bohr model
- Quantum mechanical model
- Atomic orbital
- Electronic configurations
- Orbital Quantum numbers
- Electrons in atoms and the Periodic Table

# Chapter 7 & 8

## • ELECTROMAGNETIC RADIATION

- subatomic particles (electron, photon, etc) have both PARTICLE and WAVE properties
- Light is electromagnetic radiation - crossed electric and magnetic waves:

Properties:

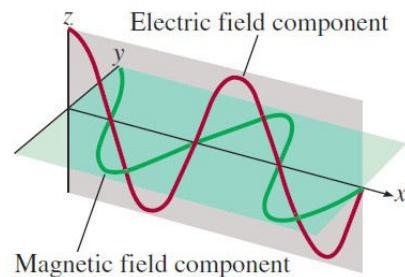
Wavelength,  $\lambda$  (nm)

Frequency,  $\nu$  ( $s^{-1}$ , Hz)

Amplitude, A

constant speed. c

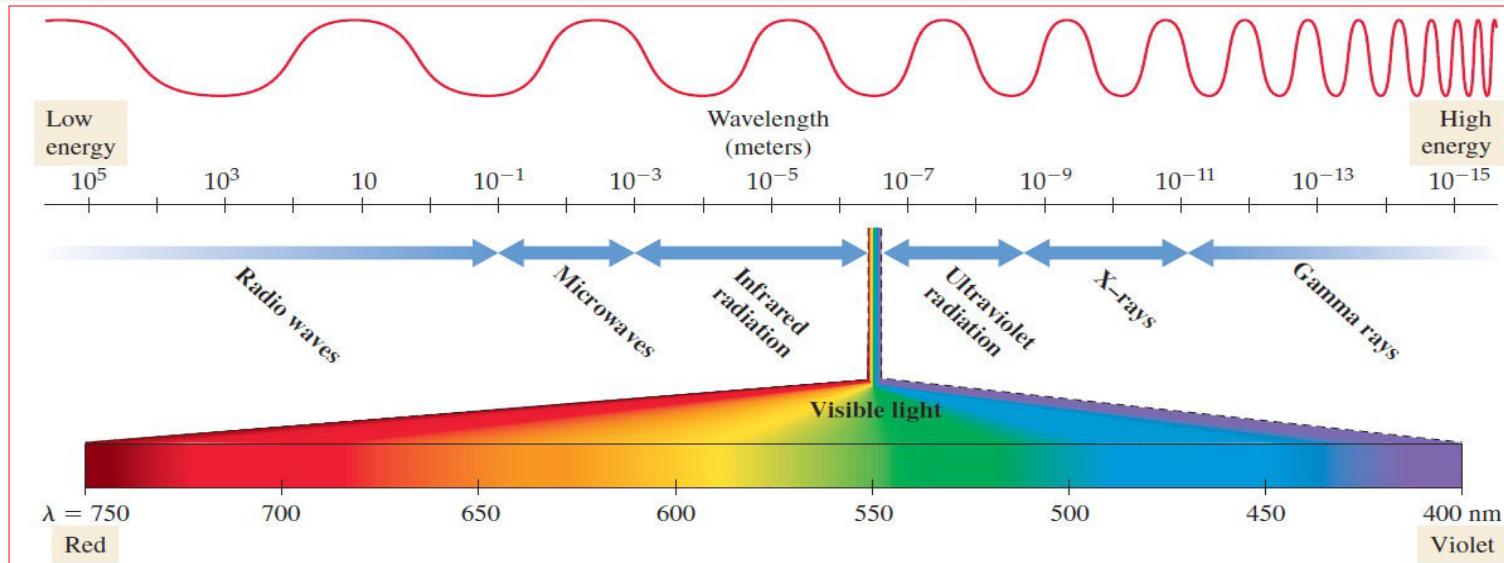
$3.00 \times 10^8 \text{ m.s}^{-1}$



- The electric field and magnetic field components of an electromagnetic wave. These two components have the same wavelength, frequency, and amplitude, but they vibrate in two mutually perpendicular planes.



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- Electromagnetic radiation ranges in wavelength from  $10^{-16} \text{ m}$  (gamma rays) to  $10^6 \text{ m}$  (radio waves). The entire range of electromagnetic radiation is called the electromagnetic spectrum.
- Electromagnetic spectrum, with short-wave length, high-frequency radiation on the right and long-wavelength, low-frequency radiation on the left.



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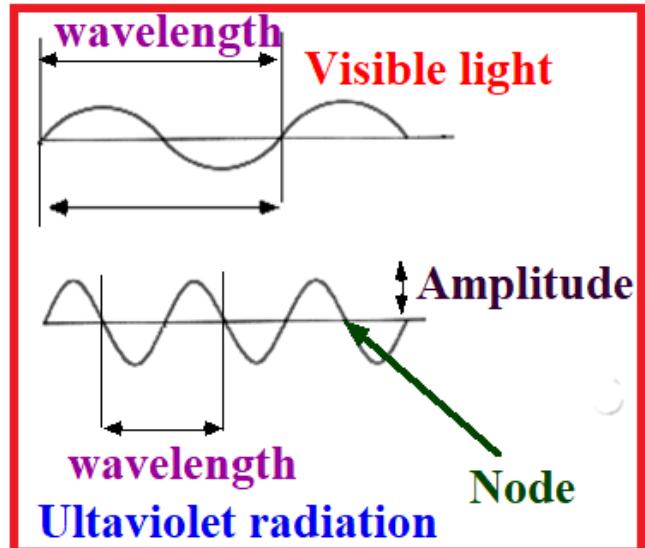
- **Example:** Arrange the three types of electromagnetic radiation—visible light, X-rays, and microwaves—in order of increasing:  
(a) wavelength (b) frequency (c) energy per photon
- **Solution:** (a) **Wavelength** Figure indicates that X-rays have the shortest wavelength, followed by visible light and then microwaves: **X-rays < visible light < microwaves.**  
(b) **Frequency** Since frequency and wavelength are inversely proportional—the longer the wavelength, the shorter the frequency—the ordering with respect to frequency is exactly the reverse of the ordering with respect to wavelength: **microwaves < visible light < X-rays.**  
(c) **Energy** per photon Energy per photon decreases with increasing wavelength but increases with increasing frequency:  
**microwaves < visible light < X-rays.**



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## • Electromagnetic Radiation

- Wavelength  $\lambda$  (lambda) is the distance between identical points on successive waves.
- The frequency  $\nu$  (nu) is the number of waves that pass through a particular point in 1 second
- Amplitude is the vertical distance from the midline of a wave to the peak or trough.





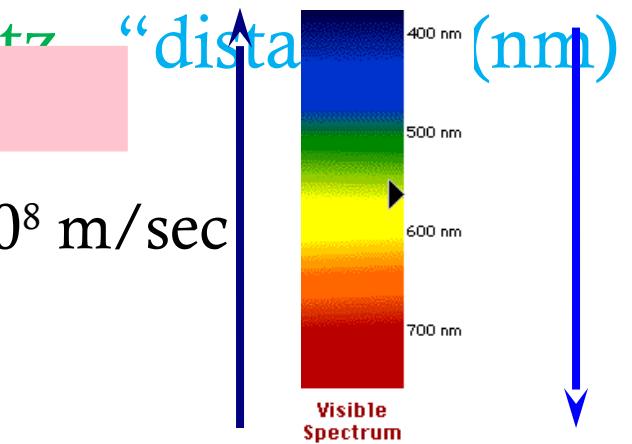
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## • **Electromagnetic Radiation**

- All waves have:  
frequency and wavelength
- symbol:  $\nu$  (Greek letter “nu”)  
“lambda”)
- units: “cycles per sec” – Hertz
- **All radiation:**  $\lambda \cdot \nu = c$

where  $c$  = velocity of light =  $3.00 \times 10^8$  m/sec

**Note: Long wavelength**  
→ small frequency  
**Short wavelength**  
→ high frequency



increasing  
frequency

increasing  
wavelength



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## • Electromagnetic Radiation

**Example:** Red light has  $\lambda = 700 \text{ nm}$ .

Calculate the frequency,  $n$ .

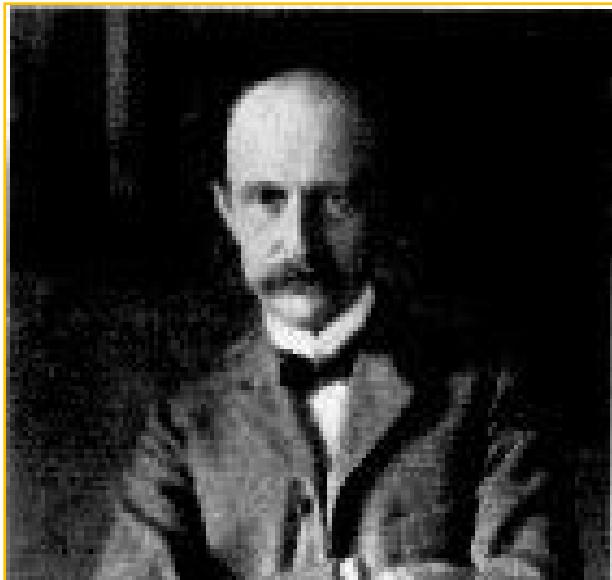
$$v = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{7.00 \times 10^{-7} \text{ m}} = 4.29 \times 10^{14} \text{ Hz}$$

- **Wave nature of light** is shown by **classical wave properties** such as
  - *interference*
  - *diffraction*



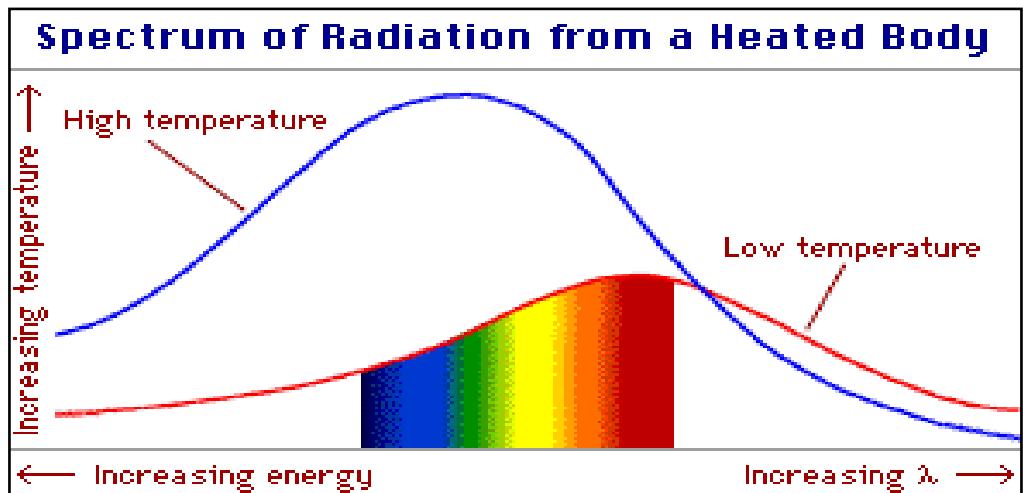
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## • Quantization of Energy



Max Planck (1858-1947)  
Solved the “ultraviolet catastrophe”

- Planck's hypothesis: An object can only gain or lose energy by absorbing or emitting radiant energy in QUANTA.





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## • Quantization of Energy

Energy of radiation is proportional to frequency.

$$E = h \cdot v$$

where  $h$  = Planck's constant =  $6.6262 \times 10^{-34}$  J·s  
 $E$ =energy,  $v$ =frequency

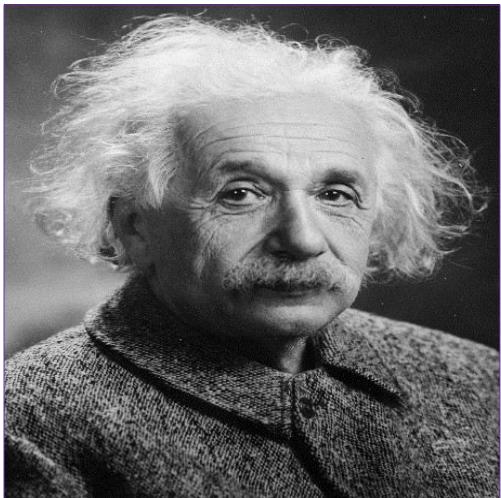
Light with large  $\lambda$  (small  $v$ ) has a small  $E$ .

Light with a short  $\lambda$  (large  $v$ ) has a large  $E$ .



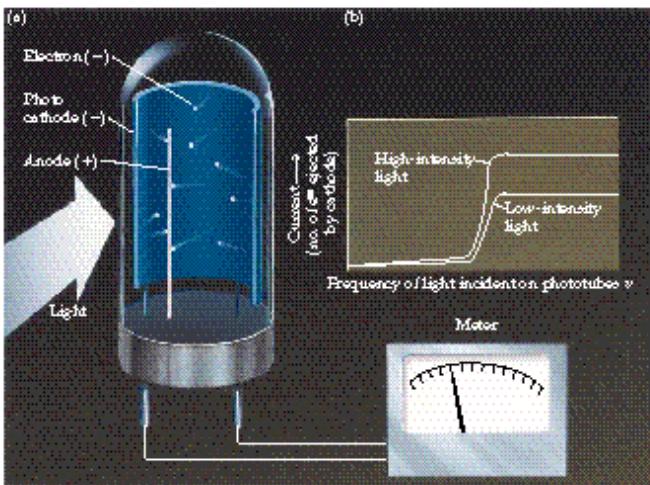
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## ❖ Photoelectric Effect



**Albert Einstein**  
**(1879-1955)**

**Photoelectric effect demonstrates the particle nature of light.**



particles of **light** are now **called photons**

**No e<sup>-</sup> observed until light of a certain minimum E is used.**

**Number of e<sup>-</sup> ejected does NOT depend on frequency, rather it depends on light intensity.**



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## ♦ Photoelectric Effect

- Classical theory said that E of ejected electron should increase with increase in light intensity — not observed!
- Experimental observations can be explained if light consists of particles called PHOTONS of discrete energy.



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## • Energy of Radiation

**PROBLEM:** Calculate the energy of 1.00 mol of photons of red light.

$$\lambda = 700 \text{ nm} \quad v = 4.29 \times 10^{14} \text{ sec}^{-1}$$

$$\begin{aligned} E &= h \cdot v \\ &= (6.63 \times 10^{-34} \text{ J}\cdot\text{s})(4.29 \times 10^{14} \text{ sec}^{-1}) \\ &= 2.85 \times 10^{-19} \text{ J per photon} \end{aligned}$$

$$\begin{aligned} E \text{ per mol} &= (2.85 \times 10^{-19} \text{ J/ph})(6.02 \times 10^{23} \text{ ph/mol}) \\ &= 171.6 \text{ kJ/mol} \end{aligned}$$

- the range of energies that can break bonds.



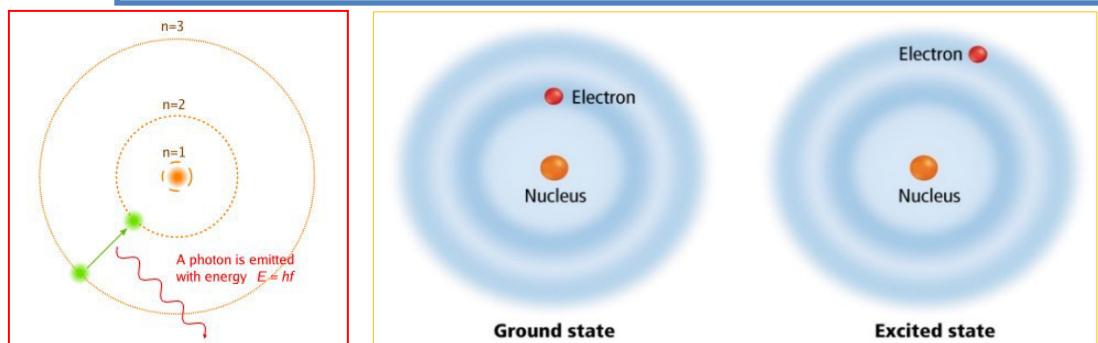
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## ◆ Bohr Model: Atomic Line Spectra



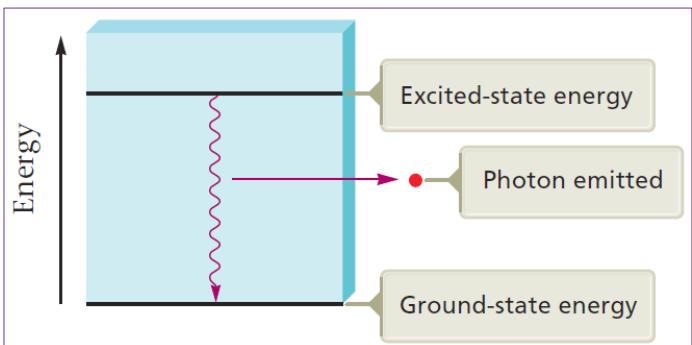
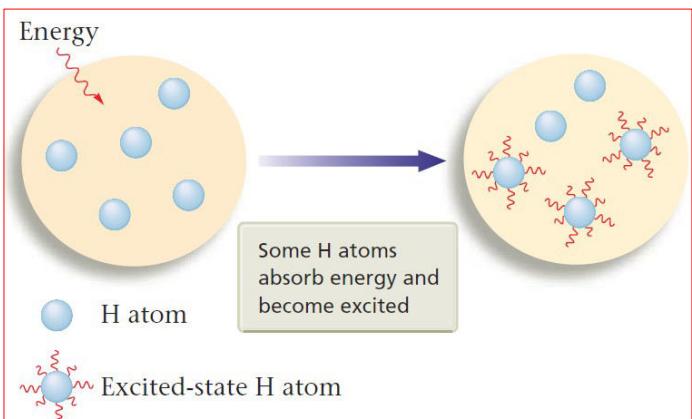
**Niels Bohr (1885-1962)**  
**(Nobel Prize, 1922)**

- **Niels Bohr (1885–1962)** was a Danish physicist and a student of Rutherford's.
- In 1913, Bohr introduced his atomic model based on the simplest atom, hydrogen (only 1 electron)
  - Bohr proposed that an electron is found only in specific circular paths, or orbits, around the nucleus.



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## ◆ Bohr Model: Atomic Line Spectra



- ◆ A quantum of energy is the amount of energy needed to move an electron from one energy level to another energy level.
- ◆ The higher the energy level, the farther it is from the nucleus.
  - Gain energy to move to higher energy levels (away from nucleus)
  - Lose energy to move to lower energy levels (closer to nucleus)

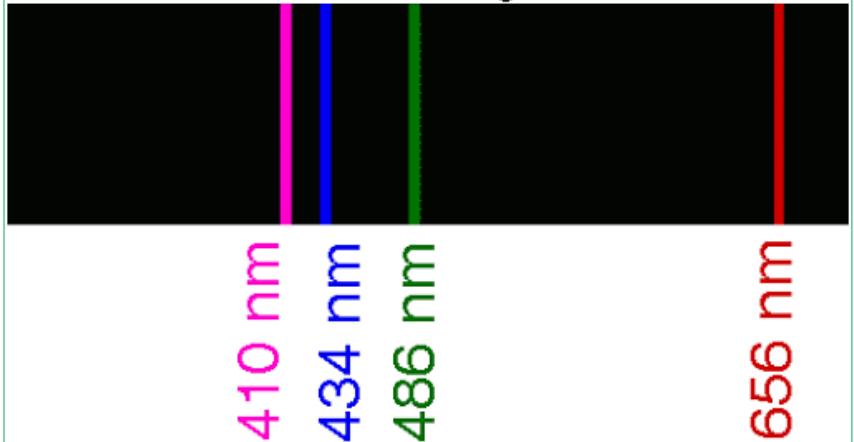


- Excited atoms emit light of only certain wavelengths
- The wavelengths of emitted light depend on the element.



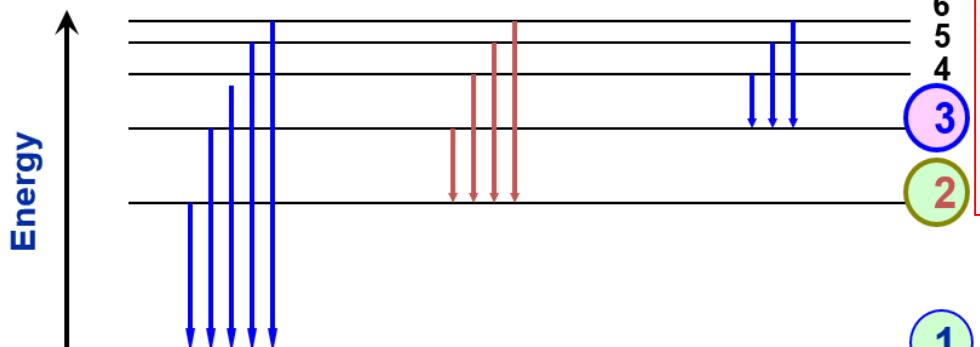
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Hydrogen emission spectrum in the visible region



Violet  
Blue  
Red

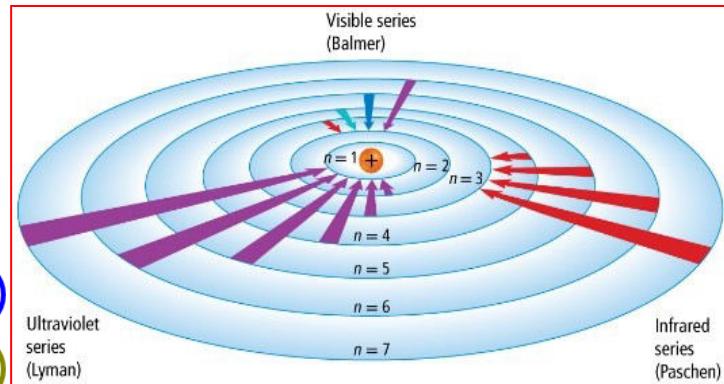
Balmer Series



$$E_n = \frac{-1312}{n^2}$$

Hydrogen atom spectra

Visible lines in H atom spectrum are called the **BALMER** series.



n

	Range	
Violet Light	380-450 nm	400 nm
Blue Light	475 nm	475 nm
Green Light	495-570 nm	510nm
Yellow Light	570 nm	570 nm
Orange Light	590 nm	590 nm
Red Light	620-750 nm	650 nm



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## From Bohr model to Quantum mechanics

- Bohr's theory was a great accomplishment and radically changed our view of matter.
- But problems existed with Bohr theory —
  - theory only successful for the H atom.
  - introduced quantum idea artificially.

So, we go on to QUANTUM or WAVE MECHANICS



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## ♦ Quantum or Wave Mechanics



L. de Broglie  
(1892-1987)

$\lambda$  for particles is called the de Broglie wavelength

- Light has both wave & particle properties
- de Broglie (1924) proposed that all moving objects have wave properties.
- For light:  $E = h\nu = hc / \lambda$
- For particles:  $E = mc^2$  (Einstein)

Therefore,  $mc = h / \lambda$

and for particles

$(\text{mass}) \times (\text{velocity}) = h / \lambda$



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## ♦ Quantum or Wave Mechanics



E. Schrodinger  
1887-1961

Schrodinger Equation

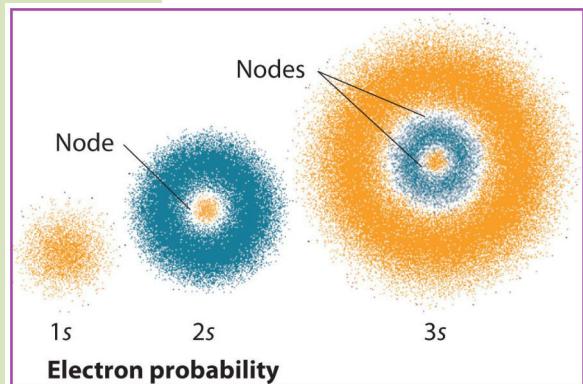
$$\hat{H} \Psi = E \Psi$$

- Rutherford's and Bohr's model focused on describing the path of the electron around the nucleus like a particle (like a small baseball).
- Austrian physicist Erwin Schrödinger (1887–1961) treated the electron as a wave.
  - The modern description of the electrons in atoms, the quantum mechanical model, comes from the mathematical solutions to the Schrödinger equation.

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## • WAVE FUNCTIONS, $\Psi$

- $\Psi$  is a function of distance and two angles.
- For 1 electron,  $\Psi$  corresponds to an ORBITAL — the region of space within which an electron is found.
- $\Psi$  does NOT describe the exact location of the electron.
- $\Psi^2$  is proportional to the probability of finding an e- at a given point.





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## • Uncertainty Principle



W. Heisenberg  
1901-1976

- Problem of defining nature of electrons in atoms solved by W. Heisenberg.
- Cannot simultaneously define the position and momentum ( $= m \cdot v$ ) of an electron.  
$$\Delta x \cdot \Delta p = h$$
- At best we can describe the position and velocity of an electron by a

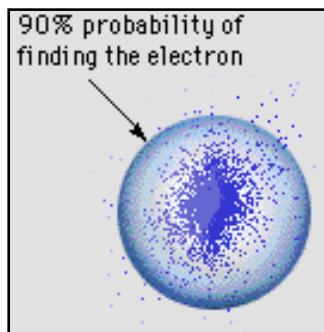
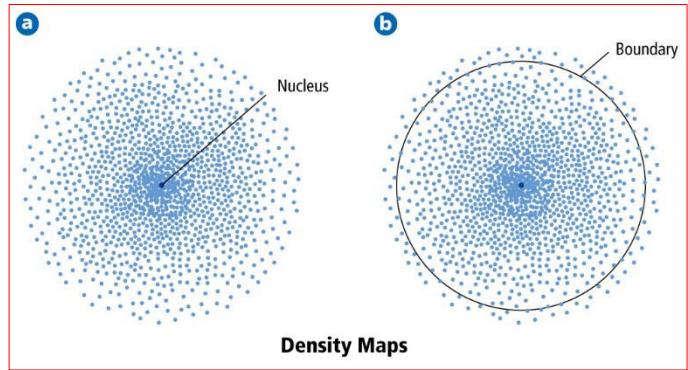
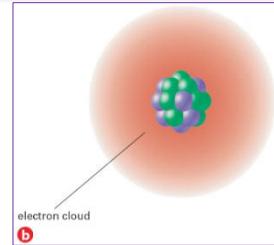
PROBABILITY DISTRIBUTION,  
which is given by  $\Psi^2$

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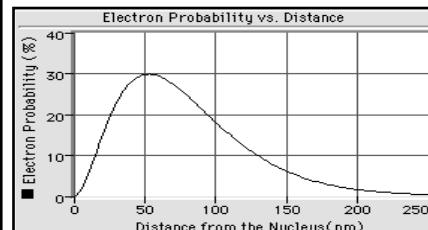
## Atomic orbital

- **Atomic orbital (fuzzy cloud)** is often thought of as a region of space in which there is a **high probability** of finding an **electron**.

- **Orbital (“electron cloud”)**
  - Region in space where there is **90% probability** of finding an  $e^-$



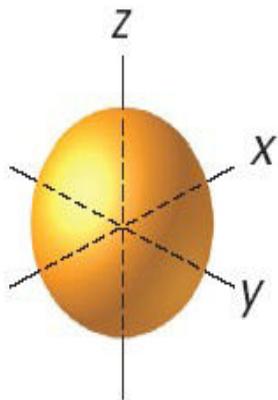
Orbital



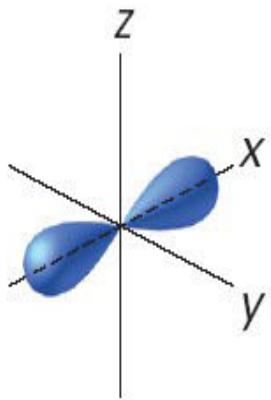
Radial Distribution Curve

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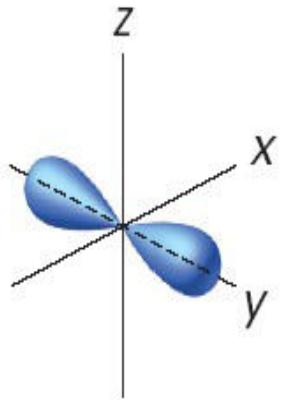
- Different atomic orbitals are denoted by letters. The **s orbitals** are spherical, and three **p orbitals** are dumbbell-shaped.



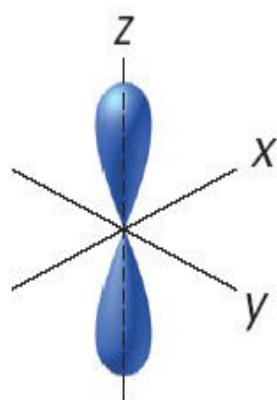
s orbital



$p_x$  orbital



$p_y$  orbital

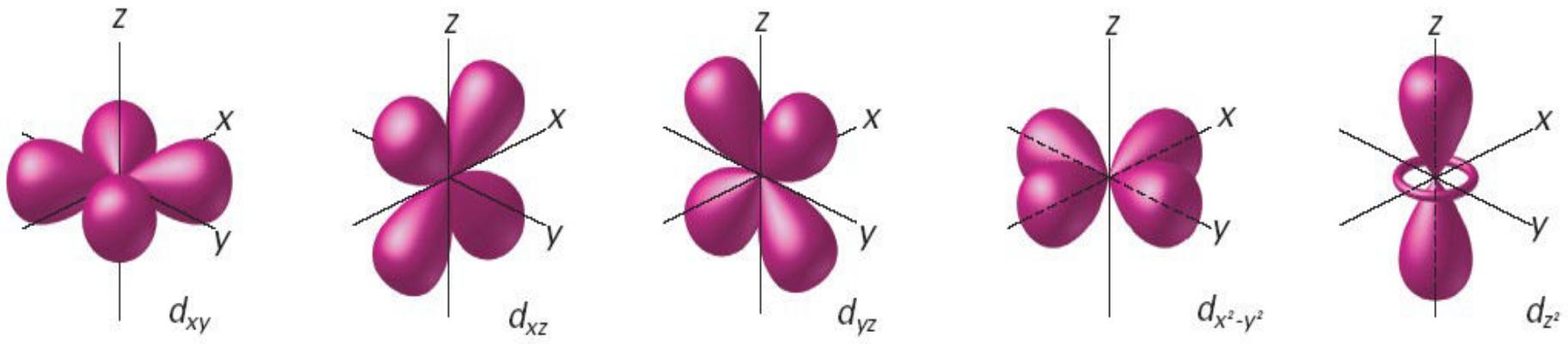


$p_z$  orbital



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- Five **d orbitals** have the same shape but different **orientations** in space.



- The **numbers** and kinds of atomic orbitals depend on the **energy sublevel**.



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- The **number of electrons** allowed in each of the **first four energy levels** are shown here.
  - A maximum of 2 electrons per orbital

Maximum Numbers of Electrons	
Energy level <i>n</i>	Maximum number of electrons
1	2
2	8
3	18
4	32

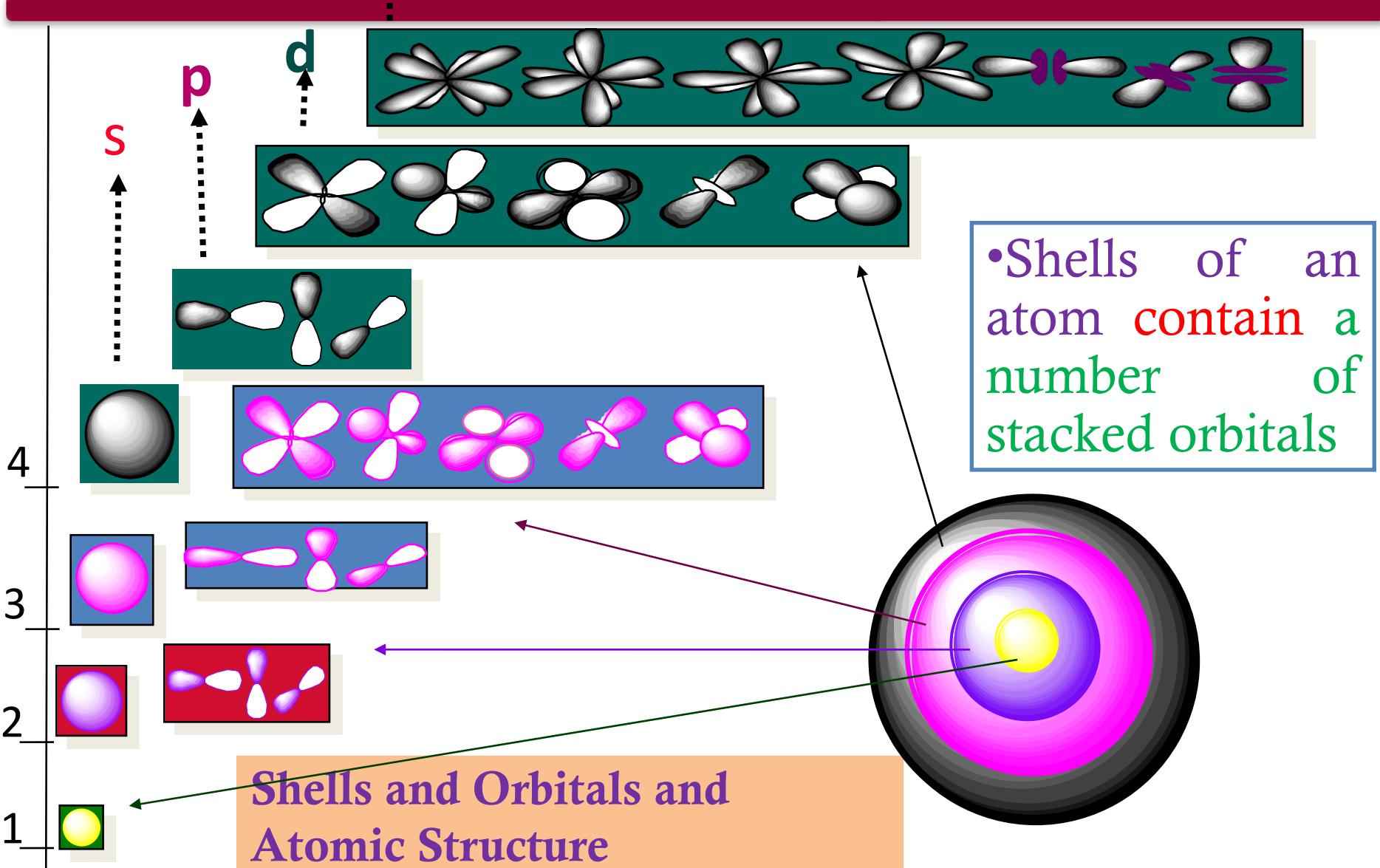
Use this to find  
the # of electrons  
in an energy level  
 $2n^2$



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Energy Level, n	# of sublevels	Letter of sublevels	# of orbitals per sublevel	# of electrons in each orbital	Total electrons in energy level
1	1	s	1	2	2
2	2	s p	1 3	2 6	8
3	3	s p d	1 3 5	2 6 10	18
4	4	s p d f	1 3 5 7	2 6 10 14	32

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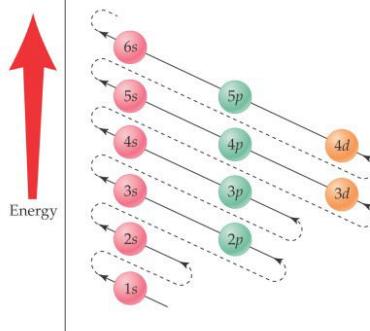
## • **Electronic Configurations**

- Electrons are arranged in various orbitals around the nuclei of atoms are called **electron configurations**.
  - Three rules—the aufbau principle, the Pauli exclusion principle, and Hund’s rule—tell you how to find the electron configurations of atoms.

### • **Aufbau Principle**

—According to the **aufbau principle**, electrons occupy the orbitals of lowest energy first.

In the ground state, the electrons will fill the atomic orbital of lowest energy.



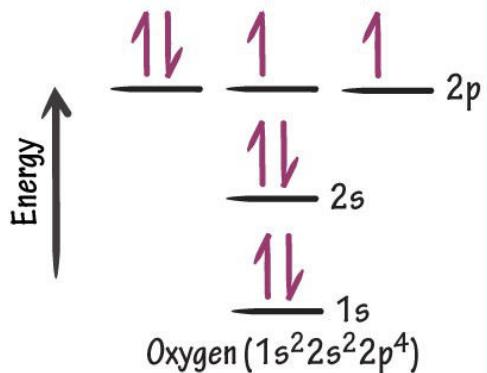


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- **Pauli Exclusion Principle**
  - According to the **Pauli exclusion principle**, no two electrons in an atom can have the same four quantum numbers.

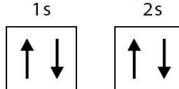
## HUND'S RULE

✓ Degenerate orbitals are occupied first by single  $e^-$  with "up" spin before being occupied by paired  $e^-$ .

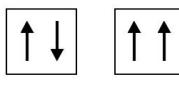


## Pauli Exclusion Principle

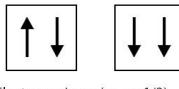
No two electrons can have the same set of quantum numbers



Correct



Incorrect



Incorrect

**Paramagnetic** substances are those that contain net unpaired spins and are attracted by a magnet.

**Diamagnetic** substances do not contain net unpaired spins and are slightly repelled by a magnet.

## Paramagnetic vs. Diamagnetic

### Paramagnetic



s orbital



p orbitals

### Diamagnetic



s orbital



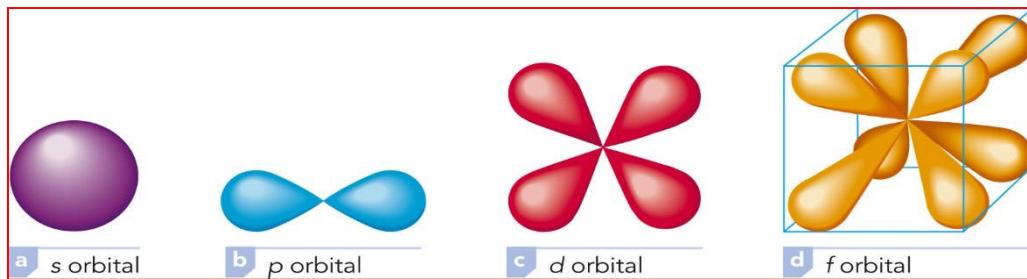
p orbitals



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## • Electron Orbitals

Subshell	Number of Orbitals
s	1
p	3
d	5
f	7





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## Electron Configurations

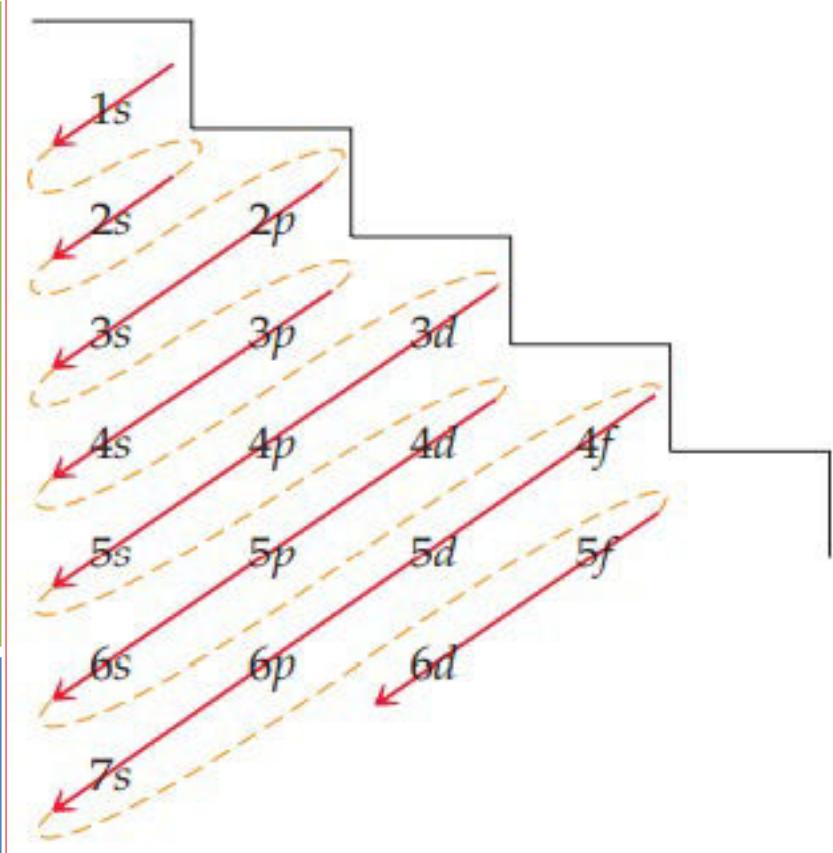
Electron Order:

1s 2s 2p 3s 3p

4s 3d 4p 5s 4d 5p

6s 4f 5d 6p 7s 5f 6d 7p

s=2; p=6; d=10; f=14





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- The electron configuration for Na is:

Na (11):  $1s^2 2s^2 2p^6 3s^1$

$Na^+$  (10):  $1s^2 2s^2 2p^6$

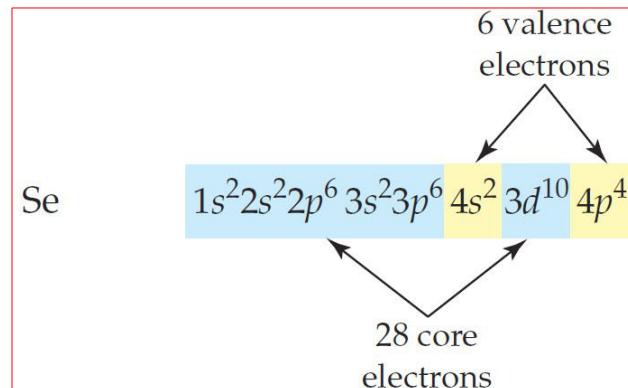
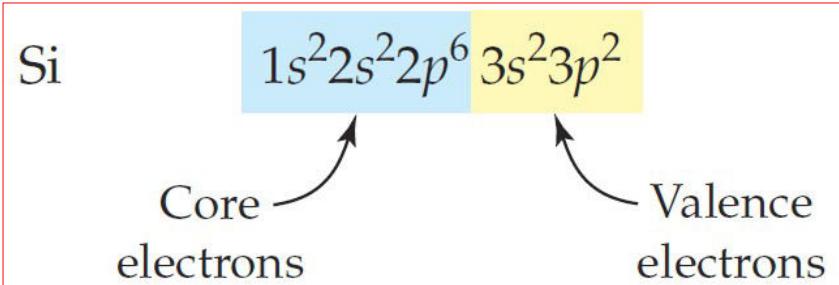
- We can abbreviate the electron configuration by indicating the innermost electrons (Core electrons) with the symbol of the preceding noble gas.
- The preceding noble gas with an atomic number less than sodium is neon, Ne. We rewrite the electron configuration:

Na: [Ne]  $3s^1$



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- Neon completes the  $2p$  subshell.
- [Ne] represents the electron configuration of neon.
- **Core electrons:** electrons in [Noble Gas].
- **Valence electrons:** electrons outside of [Noble Gas].





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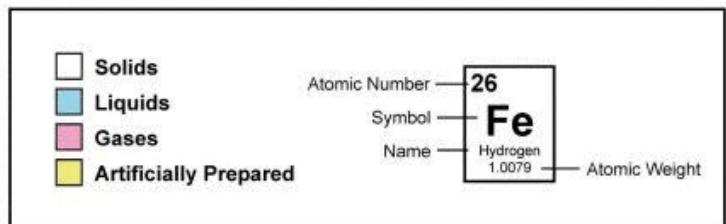
		1s	2s	2p	3s	
Li	3	<span style="color:red">1↓</span>	<span style="color:red">1</span>			$1s^2 2s^1$
Be	4	<span style="color:red">1↓</span>	<span style="color:red">1↓</span>			$1s^2 2s^2$
B	5	<span style="color:red">1↓</span>	<span style="color:red">1↓</span>	<span style="color:red">1</span>		$1s^2 2s^2 2p^1$
C	6	<span style="color:red">1↓</span>	<span style="color:red">1↓</span>	<span style="color:red">1</span> <span style="color:red">1</span>		$1s^2 2s^2 2p^2$
N	7	<span style="color:red">1↓</span>	<span style="color:red">1↓</span>	<span style="color:red">1</span> <span style="color:red">1</span> <span style="color:red">1</span>		$1s^2 2s^2 2p^3$
NE	10	<span style="color:red">1↓</span>	<span style="color:red">1↓</span>	<span style="color:red">1↓</span> <span style="color:red">1↓</span> <span style="color:red">1↓</span>		$1s^2 2s^2 2p^6$
Na	11	<span style="color:red">1↓</span>	<span style="color:red">1↓</span>	<span style="color:red">1↓</span> <span style="color:red">1↓</span> <span style="color:red">1↓</span>	<span style="color:red">1</span>	$1s^2 2s^2 2p^6 3s^1$



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## Periodic Table of the Elements

GROUP		VIII																	
IA		He																	
PERIOD	1	H	Hydrogen 1.00794																
	2	Li	Be	Lithium 6.941      Beryllium 9.01218															
	3	Na	Mg	Sodium 22.98977      Magnesium 24.3050															
	4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
	5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
	6	Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
	7	Fr	Ra		104	105	106	107	108	109	110	Uun	111	112					
				La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
				Lanthanum 138.9055	Cerium 140.116	Praseodymium 140.90765	Neodymium 144.24	Promethium (145)	Samarium 150.36	Europium 151.964	Gadolinium 157.25	Terbium 158.92534	Dysprosium 162.50	Holmium 164.93032	Erbium 167.26	Thulium 168.93421	Ytterbium 173.04	Lutetium 174.967	
				Actinium (227)	Thorium 232.0381	Protactinium 231.03588	Uranium 238.0289	Neptunium (237)	Plutonium (244)	Americium (243)	Curium (247)	Berkelium (247)	Californium (251)	Einsteinium (252)	Fermium (257)	Mendelevium (256)	Nobelium (259)	Lawrencium (262)	





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1	H	$1s^1$	20	Ca	$[Ar]4s^2$
2	He	$1s^2$	21	Sc	$[Ar]4s^2 3d^1$
3	Li	$[He]2s^1$	22	Ti	$[Ar]4s^2 3d^2$
4	Be	$[He]2s^2$	23	V	$[Ar]4s^2 3d^3$
5	B	$[He]2s^2 2p^1$	24	Cr	$[Ar]4s^2 3d^4$
6	C	$[He]2s^2 2p^2$	25	Mn	$[Ar]4s^2 3d^5$
7	N	$[He]2s^2 2p^3$	26	Fe	$[Ar]4s^2 3d^6$
8	O	$[He]2s^2 2p^4$	27	Co	$[Ar]4s^2 3d^7$
9	F	$[He]2s^2 2p^5$	28	Ni	$[Ar]4s^2 3d^8$
10	Ne	$[He]2s^2 2p^6$	29	Cu	$[Ar]4s^1 3d^{10}$
11	Na	$[Ne]3s^1$	38	Zn	$[Ar]4s^2 3d^10$
12	Mg	$[Ne]3s^2$	31	Ga	$[Ar]4s^2 3d^10 4p^1$
13	Al	$[Ne]3s^2 3p^1$	32	Ge	$[Ar]4s^2 3d^10 4p^2$
14	Si	$[Ne]3s^2 3p^2$	33	As	$[Ar]4s^2 3d^10 4p^3$
15	P	$[Ne]3s^2 3p^3$	34	Se	$[Ar]4s^2 3d^10 4p^4$
16	S	$[Ne]3s^2 3p^4$	35	Br	$[Ar]4s^2 3d^10 4p^5$
17	Cl	$[Ne]3s^2 3p^5$	36	Kr	$[Ar]4s^2 3d^10 4p^6$
18	Ar	$[Ne]3s^2 3p^6$	37	Rb	$[Kr]5s^1$
19	K	$[Ar]4s^1$	38	Sr	$[Kr]5s^2$



# Chapter 7 & 8

39	Y	[Kr]5s <sup>2</sup> 4d <sup>1</sup>	58	Ce	[Xe]6s <sup>2</sup> 4f <sup>1</sup> 5d <sup>1</sup>
40	Zr	[Kr]5s <sup>2</sup> 4d <sup>2</sup>	59	Pr	[Xe]6s <sup>2</sup> 4f <sup>2</sup>
41	Nb	[Kr]5s <sup>1</sup> 4d <sup>4</sup>	60	Nd	[Xe]6s <sup>2</sup> 4f <sup>3</sup>
42	Mo	[Kr]5s <sup>1</sup> 4d <sup>5</sup>	61	Pm	[Xe]6s <sup>2</sup> 4f <sup>4</sup>
43	Tc	[Kr]5s <sup>1</sup> 4d <sup>6</sup>	62	Sm	[Xe]6s <sup>2</sup> 4f <sup>5</sup>
44	Ru	[Kr]5s <sup>1</sup> 4d <sup>7</sup>	63	Eu	[Xe]6s <sup>2</sup> 4f <sup>6</sup>
45	Rh	[Kr]5s <sup>1</sup> 4d <sup>8</sup>	64	Gd	[Xe]6s <sup>2</sup> 4f <sup>7</sup> 5d <sup>1</sup>
46	Pd	[Kr]4d <sup>10</sup>	65	Tb	[Xe]6s <sup>2</sup> 4f <sup>8</sup>
47	Ag	[Kr]5s <sup>1</sup> 4d <sup>10</sup>	66	Dy	[Xe]6s <sup>2</sup> 4f <sup>9</sup>
48	Cd	[Kr]5s <sup>2</sup> 4d <sup>10</sup>	67	Ho	[Xe]6s <sup>2</sup> 4f <sup>10</sup>
49	In	[Kr]5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>1</sup>	68	Er	[Xe]6s <sup>2</sup> 4f <sup>11</sup>
50	Sn	[Kr]5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>2</sup>	69	Tm	[Xe]6s <sup>2</sup> 4f <sup>12</sup>
51	Sb	[Kr]5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>3</sup>	70	Yb	[Xe]6s <sup>2</sup> 4f <sup>13</sup>
52	Te	[Kr]5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>4</sup>	71	Lu	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>1</sup>
53	I	[Kr]5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>5</sup>	72	Hf	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>2</sup>
54	Xe	[Kr]5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>6</sup>	73	Ta	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>3</sup>
55	Cs	[Xe]6s <sup>1</sup>	74	W	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>4</sup>
56	Ba	[Xe]6s <sup>2</sup>	75	Re	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>5</sup>
57	La	[Xe]6s <sup>2</sup> 5d <sup>1</sup>	76	Os	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>6</sup>



# Chapter 7 & 8

77	Ir	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>9</sup>	95	Am	[Rn]7s <sup>2</sup> 5f <sup>7</sup>
78	Pt	[Xe]6s <sup>1</sup> 4f <sup>14</sup> 5d <sup>9</sup>	96	Cm	[Rn]7s <sup>2</sup> 5f <sup>7</sup> 6d <sup>1</sup>
79	Au	[Xe]6s <sup>1</sup> 4f <sup>14</sup> 5d <sup>10</sup>	97	Bk	[Rn]7s <sup>2</sup> 5f <sup>9</sup>
80	Hg	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup>	98	Cf	[Rn]7s <sup>2</sup> 5f <sup>10</sup>
81	Tl	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup> 6p <sup>1</sup>	99	Es	[Rn]7s <sup>2</sup> 5f <sup>11</sup>
82	Pb	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup> 6p <sup>2</sup>	100	Fm	[Rn]7s <sup>2</sup> 5f <sup>12</sup>
83	Bi	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup> 6p <sup>3</sup>	101	Md	[Rn]7s <sup>2</sup> 5f <sup>13</sup>
84	Po	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup> 6p <sup>4</sup>	102	No	[Rn]7s <sup>2</sup> 5f <sup>14</sup>
85	At	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup> 6p <sup>5</sup>	103	Lr	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>1</sup>
86	Rn	[Xe]6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>10</sup> 6p <sup>6</sup>	104	[104]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>2</sup>
87	Fr	[Rn]7s <sup>1</sup>	105	[105]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>2</sup>
88	Ra	[Rn]7s <sup>2</sup>	106	[106]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>4</sup>
89	Ac	[Rn]7s <sup>2</sup> 6d <sup>1</sup>	107	[107]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>5</sup>
90	Th	[Rn]7s <sup>2</sup> 6d <sup>2</sup>	108	[108]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>6</sup>
91	Pa	[Rn]7s <sup>2</sup> 5f <sup>2</sup> 6d <sup>1</sup>	109	[109]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>7</sup>
92	U	[Rn]7s <sup>2</sup> 5f <sup>3</sup> 6d <sup>1</sup>	110	[110]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>8</sup>
93	Np	[Rn]7s <sup>2</sup> 5f <sup>4</sup> 6d <sup>1</sup>	111	[111]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>9</sup>
94	Pu	[Rn]7s <sup>2</sup> 5f <sup>6</sup>	112	[112]	[Rn]7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>10</sup>



# Chapter 7 & 8

## Example:

- Write the ground-state electron configurations for (a) sulfur (S) and (b) palladium (Pd), which is diamagnetic.

### Answer:

(a) Sulfur (S) has 16 electrons:  $1s^2 2s^2 2p^6 3s^2 3p^4$

(b) Palladium (Pd) has 46 electrons:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$

- Because palladium is diamagnetic, all the electrons are paired



# Chapter 7 & 8

## Example:

- Give the condensed and full electron configurations for the following elements:  
**(a) Technetium (Tc;  $Z = 43$ ) (b) Lead (Pb;  $Z = 82$ )**
- Answer:

**(a) For Tc ( $Z = 43$ )**

**Condensed:** [Kr]5s<sup>2</sup>4d<sup>5</sup>

**Full:**

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^5$

**(b) For Pb ( $Z = 82$ )**

**Condensed:** [Xe] 6s<sup>2</sup>4f<sup>4</sup>5d<sup>10</sup>6p<sup>2</sup>

**Full:**

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^4 5d^{10} 6p^2$



# Chapter 7 & 8

	Periodic Table of Elements																																			
	1A		2A																8A																	
Core	1	H 1s <sup>1</sup>	2																18																	
[He]	3	Li 2s <sup>1</sup>	4	Be 2s <sup>2</sup>																																
[Ne]	11	Na 3s <sup>1</sup>	12	Mg 3s <sup>2</sup>	3B 3	4B 4	5B 5	6B 6	7B 7	8B		8 9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17																
[Ar]	19	K 4s <sup>1</sup>	20	Ca 4s <sup>2</sup>	21	Sc 3d <sup>1</sup> 4s <sup>2</sup>	22	Ti 3d <sup>2</sup> 4s <sup>2</sup>	23	V 3d <sup>3</sup> 4s <sup>2</sup>	24	Cr 3d <sup>5</sup> 4s <sup>1</sup>	25	Mn 3d <sup>5</sup> 4s <sup>2</sup>	26	Fe 3d <sup>6</sup> 4s <sup>2</sup>	27	Co 3d <sup>7</sup> 4s <sup>2</sup>	28	Ni 3d <sup>8</sup> 4s <sup>2</sup>	29	Cu 3d <sup>10</sup> 4s <sup>1</sup>	30	Zn 3d <sup>10</sup> 4s <sup>2</sup>	31	Ga 3d <sup>10</sup> 4s <sup>2</sup>	32	Ge 3d <sup>10</sup> 4s <sup>2</sup>	33	As 3d <sup>10</sup> 4s <sup>2</sup>	34	Se 3d <sup>10</sup> 4s <sup>2</sup>	35	Br 3d <sup>10</sup> 4s <sup>2</sup>	36	Kr 3d <sup>10</sup> 4s <sup>2</sup>
[Kr]	37	Rb 5s <sup>1</sup>	38	Sr 5s <sup>2</sup>	39	Y 4d <sup>1</sup> 5s <sup>2</sup>	40	Zr 4d <sup>2</sup> 5s <sup>2</sup>	41	Nb 4d <sup>3</sup> 5s <sup>2</sup>	42	Mo 4d <sup>5</sup> 5s <sup>1</sup>	43	Tc 4d <sup>5</sup> 5s <sup>2</sup>	44	Ru 4d <sup>7</sup> 5s <sup>1</sup>	45	Rh 4d <sup>8</sup> 5s <sup>1</sup>	46	Pd 4d <sup>10</sup>	47	Ag 4d <sup>10</sup> 5s <sup>1</sup>	48	Cd 4d <sup>10</sup> 5s <sup>2</sup>	49	In 4d <sup>10</sup> 5s <sup>2</sup>	50	Sn 4d <sup>10</sup> 5s <sup>2</sup>	51	Sb 4d <sup>10</sup> 5s <sup>2</sup>	52	Te 4d <sup>10</sup> 5s <sup>2</sup>	53	I 4d <sup>10</sup> 5s <sup>2</sup>	54	Xe 4d <sup>10</sup> 5s <sup>2</sup>
[Xe]	55	Cs 6s <sup>1</sup>	56	Ba 6s <sup>2</sup>	71	Lu 4f <sup>14</sup> 5d <sup>1</sup> 6s <sup>2</sup>	72	Hf 4f <sup>14</sup> 5d <sup>2</sup> 6s <sup>2</sup>	73	Ta 4f <sup>14</sup> 5d <sup>3</sup> 6s <sup>2</sup>	74	W 4f <sup>14</sup> 5d <sup>4</sup> 6s <sup>2</sup>	75	Re 4f <sup>14</sup> 5d <sup>5</sup> 6s <sup>2</sup>	76	Os 4f <sup>14</sup> 5d <sup>6</sup> 6s <sup>2</sup>	77	Ir 4f <sup>14</sup> 5d <sup>7</sup> 6s <sup>2</sup>	78	Pt 4f <sup>14</sup> 5d <sup>9</sup> 6s <sup>1</sup>	79	Au 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>1</sup>	80	Hg 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>	81	Tl 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>	82	Pb 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>	83	Bi 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>	84	Po 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>	85	At 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>	86	Rn 4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup>
[Rn]	87	Fr 7s <sup>1</sup>	88	Ra 7s <sup>2</sup>	103	Lr 5f <sup>14</sup> 6d <sup>1</sup> 7s <sup>2</sup>	104	Rf 5f <sup>14</sup> 6d <sup>2</sup> 7s <sup>2</sup>	105	Db 5f <sup>14</sup> 6d <sup>3</sup> 7s <sup>2</sup>	106	Sg 5f <sup>14</sup> 6d <sup>4</sup> 7s <sup>2</sup>	107	Bh 5f <sup>14</sup> 6d <sup>5</sup> 7s <sup>2</sup>	108	Hs 5f <sup>14</sup> 6d <sup>6</sup> 7s <sup>2</sup>	109	Mt 5f <sup>14</sup> 6d <sup>7</sup> 7s <sup>2</sup>	110		111		112		114		116									
[Xe]	Lanthanide series		Actinide series																70																	
[Rn]	Metals																Metalloids		Nonmetals																	
	57	La 5d <sup>1</sup> 6s <sup>2</sup>	58	Ce 4f <sup>1</sup> 5d <sup>1</sup> 6s <sup>2</sup>	59	Pr 4f <sup>2</sup> 6s <sup>2</sup>	60	Nd 4f <sup>3</sup> 6s <sup>2</sup>	61	Pm 4f <sup>5</sup> 6s <sup>2</sup>	62	Sm 4f <sup>6</sup> 6s <sup>2</sup>	63	Eu 4f <sup>7</sup> 6s <sup>2</sup>	64	Gd 4f <sup>9</sup> 6s <sup>2</sup>	65	Tb 4f <sup>10</sup> 6s <sup>2</sup>	66	Dy 4f <sup>11</sup> 6s <sup>2</sup>	67	Ho 4f <sup>12</sup> 6s <sup>2</sup>	68	Er 4f <sup>13</sup> 6s <sup>2</sup>	69	Tm 4f <sup>14</sup> 6s <sup>2</sup>	70	Yb 4f <sup>14</sup> 6s <sup>2</sup>								
	89	Ac 6d <sup>1</sup> 7s <sup>2</sup>	90	Th 6d <sup>2</sup> 7s <sup>2</sup>	91	Pa 5f <sup>2</sup> 6d <sup>1</sup> 7s <sup>2</sup>	92	U 5f <sup>3</sup> 6d <sup>1</sup> 7s <sup>2</sup>	93	Np 5f <sup>4</sup> 6d <sup>1</sup> 7s <sup>2</sup>	94	Pu 5f <sup>6</sup> 7s <sup>2</sup>	95	Am 5f <sup>7</sup> 7s <sup>2</sup>	96	Cm 5f <sup>9</sup> 7s <sup>2</sup>	97	Bk 5f <sup>10</sup> 7s <sup>2</sup>	98	Cf 5f <sup>11</sup> 7s <sup>2</sup>	99	Es 5f <sup>12</sup> 7s <sup>2</sup>	100	Fm 5f <sup>13</sup> 7s <sup>2</sup>	101	Md 5f <sup>14</sup> 7s <sup>2</sup>	102	No 5f <sup>14</sup> 7s <sup>2</sup>								



# Chapter 7 & 8

## • Orbital Quantum Numbers

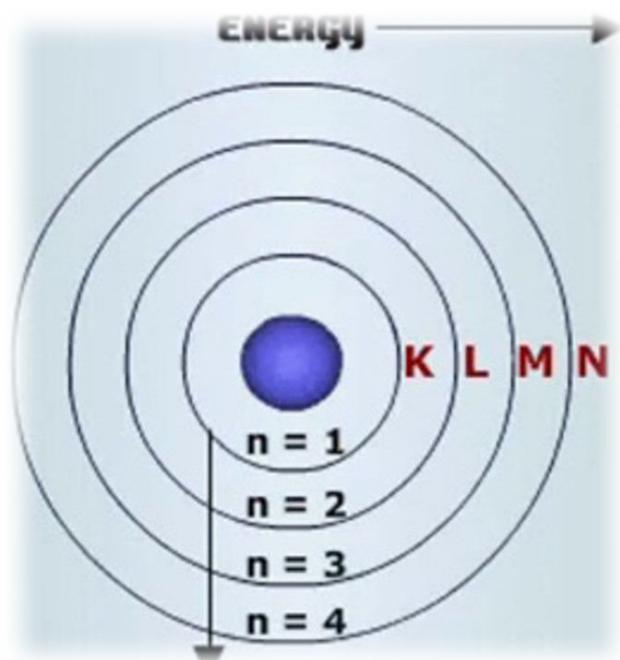
- An atomic orbital is defined by 4 quantum numbers:
  - $n$     $l$     $m_l$     $m_s$
- Electrons are arranged in shells and subshells of ORBITALS.
  - $n$  → shell
  - $l$  → subshell
  - $m_l$  → designates an orbital within a subshell
  - $m_s$  → orientation of rotation of electron around itself



# Chapter 7 & 8

1. Principal Quantum Number
2. Orbital Quantum Number
3. Magnetic Quantum Number
4. Spin Quantum Number

- Principal Quantum Number, n:
- Indicates main energy levels  $n = 1, 2, 3, 4\dots$
- Each main energy level has sub-levels
- Max # electrons =  $2n^2$





# Chapter 7 & 8

- Orbital Quantum Number,  $\ell$   
(Angular Momentum Quantum Number)

- Indicates shape of orbital sublevels

- $\ell = 0$  to  $(n-1)$

$\ell$  sublevel

0 s

1 p

2 d

3 f

4 g



# Chapter 7 & 8

## • Magnetic Quantum Number, $m_l$

- Indicates the orientation of the orbital in space.
- Values of  $m_l$ : integers -1 to 1
- The number of values represents the number of orbitals.
- Example:

for  $l=0, 1, 2$ ,  $m_l = -2, -1, 0, +1, +2$

Which sublevel does this represent?

Answer: d



# Chapter 7 & 8

**# of Orbital =  $n^2$**

**# of Orbital =  $2n^2$**

$n$	$l$	$m_l^*$	Subshell Notation	Number of Orbitals in the Subshell
1	0	0	1s	1
2	0	0	2s	1
2	1	-1, 0, +1	2p	3
3	0	0	3s	1
3	1	-1, 0, +1	3p	3
3	2	-2, -1, 0, +1, +2	3d	5
4	0	0	4s	1
4	1	-1, 0, +1	4p	3
4	2	-2, -1, 0, +1, +2	4d	5
4	3	-3, -2, -1, 0, +1, +2, +3	4f	7



# Chapter 7 & 8

**Problem:** What values of the angular momentum ( $l$ ) and magnetic ( $m_l$ ) quantum numbers are allowed for a principal quantum number ( $n$ ) of 3? How many orbitals are allowed for  $n = 3$ ?

$$n = 3, l = 0, m_l = 0 \checkmark$$

$$l = 1, m_l = -1, 0, +1 \checkmark$$

$$l = 2, m_l = -2, -1, 0, +1, +2 \checkmark$$

Total # orbitals for  $n=3$  is 9

For any  $n$ , # of orbitals is  $n^2$

For any  $n$ , # of electrons is  $2n^2$



# Chapter 7 & 8

**Problem:** Give the name, magnetic quantum numbers, and number of orbitals for each sublevel with the following quantum numbers:

- (a)  $n = 3, l = 2$  (b)  $n = 2, l = 0$  (c)  $n = 5, l = 1$  (d)  $n = 4, l = 3$

$n=3, l=2$	<u>Name</u> 3d	$m_l$ $-2, -1, 0, +1, +2$	<u># orbitals</u> 5
$n=2, l=0$	2s	0	1
$n=5, l=1$	5p	$-1, 0, +1$	3
$n=4, l=3$	4f	$-3, -2, -1, 0, +1, +2, +3$	7



# Chapter 7 & 8

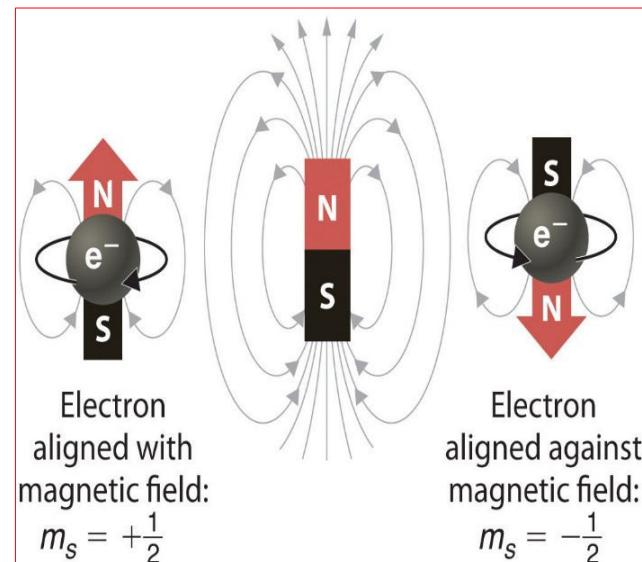
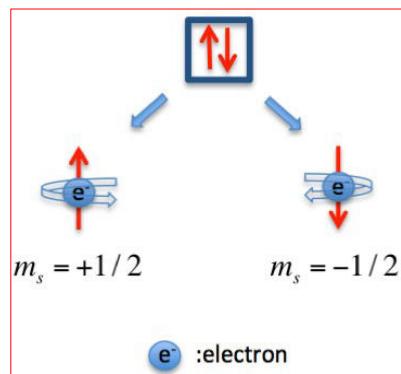
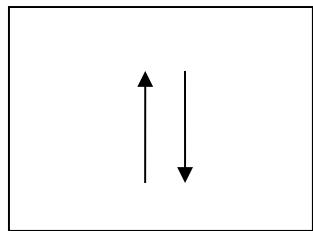
$n$	$\ell$ 0 to $(n-1)$	$m_\ell$ $(-\ell, \dots, +\ell)$	No. of orbitals	No. of electrons
1	0 ( <i>s</i> )	0	1	2
2	0 ( <i>s</i> )	0	1	2
	1 ( <i>p</i> )	-1, 0, 1	3	6
3	0 ( <i>s</i> )	0	1	2
	1 ( <i>p</i> )	-1, 0, 1	3	6
	2 ( <i>d</i> )	-2, 1, 0, 1, 2	5	10
4	0 ( <i>s</i> )	0	1	2
	1 ( <i>p</i> )	-1, 0, 1	3	6
	2 ( <i>d</i> )	-2, 1, 0, 1, 2	5	10
	3 ( <i>f</i> )	-3, -2, 1, 0, 1, 2, 3	7	14
				$2n^2$



# Chapter 7 & 8

## • Electron Spin Quantum Number, ( $m_s$ or $s$ )

- Indicates the spin of the electron (clockwise or counterclockwise).
- Values of  $m_s$ :  $+1/2, -1/2$





# Chapter 7 & 8

<b><i>n</i></b>	<b><i>ℓ</i></b>	<b><i>m<sub>ℓ</sub></i></b>	<b>Number of Orbitals</b>	<b>Atomic Orbital Designations</b>
1	0	0	1	1s
2	0	0	1	2s
	1	-1, 0, 1	3	2p <sub>x</sub> , 2p <sub>y</sub> , 2p <sub>z</sub>
3	0	0	1	3s
	1	-1, 0, 1	3	3p <sub>x</sub> , 3p <sub>y</sub> , 3p <sub>z</sub>
	2	-2, -1, 0, 1, 2	5	3d <sub>xy</sub> , 3d <sub>yz</sub> , 3d <sub>xz</sub> , 3d <sub>x<sup>2</sup>-y<sup>2</sup></sub> , 3d <sub>z<sup>2</sup></sub>
:	:	:	:	:
:	:	:	:	:
:	:	:	:	:



# Chapter 7 & 8

## Example:

- List the values of the four quantum numbers for orbitals in the 3d sublevel.
- Answer:

$n=3$

$l = 0 \text{ to } (n-1) = 0, 1, 2$

$m_l = -1 \text{ to } +1 = -2, -1, 0, +1, +2$

$m_s = +1/2, -1/2$  for each pair of electrons



# Chapter 7 & 8

## Example:

- List the values of  $n$ ,  $l$ , and  $m_l$  for orbitals in the 4d subshell.
- **Answer:** As we saw earlier, the number given in the designation of the subshell is the principal quantum number, so in this case  $n = 4$ . The letter designates (d) the type of orbital. Because we are dealing with d orbitals,  $l = 0$  to  $(n-1) = 0, 1, 2, 3$ . The values of  $m_l$  can vary from  $-1$  to  $+1$ . Therefore,  $m_l$  can be  $-3, -2, -1, 0, +1, +2, +3$ .



# Chapter 7 & 8

## p Orbitals

For  $n = 2$ ,  $l = 0$  to  $(n-1) = 0$  and  $1$

There are 2 types of orbitals – 2 subshells

For  $l = 0$ ,  $m_l = 0$

this is a s subshell

For  $l = 1$ ,  $m_l = -1$  to  $+1 = -1, 0, +1$

this is a p subshell with 3 orbitals



# Chapter 7 & 8

For  $n = 3$ , what are the values of  $l$ ?

$l = 0$  to  $(n-1) = 0, 1, 2$

and so there are 3 subshells in the shell.

## d Orbitals

For  $l = 0, m_l = 0$

→ s subshell with single orbital

For  $l = 1, m_l = -1, 0, +1$

→ p subshell with 3 orbitals

For  $l = 2, m_l = -2, -1, 0, +1, +2$

→ d subshell with 5 orbitals



# Chapter 7 & 8

## •Factors Affecting Atomic Size

- The size of an atom is affected mainly by two factors:
  - 1) Positive Charge in the Nucleus ( $Z$ )
  - 2) Number of Shells ( $n$  value) around the Nucleus
- A *higher positive charge* pulls the electrons closer to the nucleus and *reduces* the atomic size.
- The more the number of shells, the greater the *shielding effect* from the inner electrons, and this *increases* the atomic size.

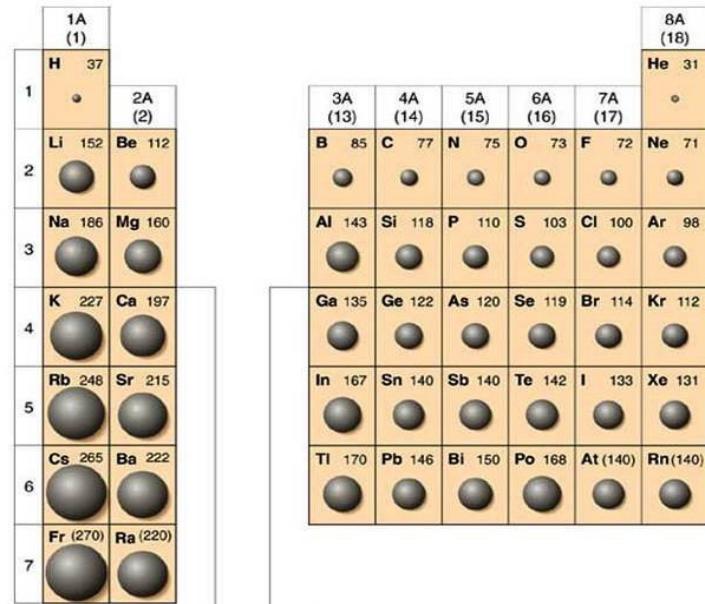


# Chapter 7 & 8

## • Trends in Atomic Size

- Atomic size ***shrinks*** across a period. The positive nuclear charge increases as we move to the right, but the **number of shells (*n* value)** remains the same.

- Atomic size ***increases*** down a group. This time we are adding a new shell (**higher *n* value**) every time we move down the group, and this **increases the shielding effect** by the inner electrons.



	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	(8)	8B (9)	(10)	1B (11)	2B (12)
4	Sc 162	Tl 147	V 134	Cr 128	Mn 127	Fe 126	Co 125	Ni 124	Cu 128	Zn 134
5	Y 180	Zr 160	Nb 146	Mo 139	Tc 136	Ru 134	Rh 134	Pd 137	Ag 144	Cd 151
6	La 187	Hf 159	Ta 146	W 139	Re 137	Os 135	Ir 136	Pt 138	Au 144	Hg 151



# Chapter 7 & 8

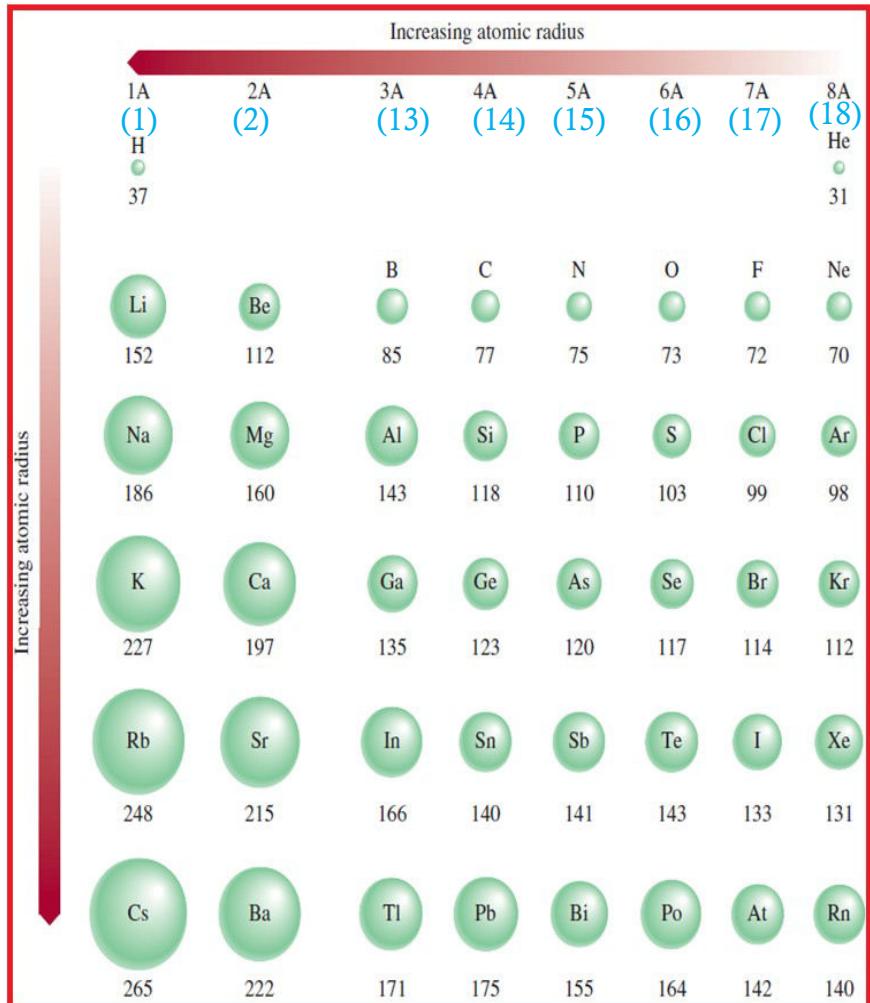
Main-Group Elements (s block)		Transition Elements (d block)															Main-Group Elements (p block)					
Period number: highest occupied energy level	1A (1)														8A (18)							
	ns <sup>1</sup>													ns <sup>2np<sup>6</sup></sup>								
1	1 H 1s <sup>1</sup>	2A (2) ns <sup>2</sup>														3A (13) ns <sup>2np<sup>1</sup></sup>	4A (14) ns <sup>2np<sup>2</sup></sup>	5A (15) ns <sup>2np<sup>3</sup></sup>	6A (16) ns <sup>2np<sup>4</sup></sup>	7A (17) ns <sup>2np<sup>5</sup></sup>	2 He 1s <sup>2</sup>	
2	3 Li 2s <sup>1</sup>	4 Be 2s <sup>2</sup>														5 B 2s <sup>2</sup> 2p <sup>1</sup>	6 C 2s <sup>2</sup> 2p <sup>2</sup>	7 N 2s <sup>2</sup> 2p <sup>3</sup>	8 O 2s <sup>2</sup> 2p <sup>4</sup>	9 F 2s <sup>2</sup> 2p <sup>5</sup>	10 Ne 2s <sup>2</sup> 2p <sup>6</sup>	
3	11 Na 3s <sup>1</sup>	12 Mg 3s <sup>2</sup>														13 Al 3s <sup>2</sup> 3p <sup>1</sup>	14 Si 3s <sup>2</sup> 3p <sup>2</sup>	15 P 3s <sup>2</sup> 3p <sup>3</sup>	16 S 3s <sup>2</sup> 3p <sup>4</sup>	17 Cl 3s <sup>2</sup> 3p <sup>5</sup>	18 Ar 3s <sup>2</sup> 3p <sup>6</sup>	
4	19 K 4s <sup>1</sup>	20 Ca 4s <sup>2</sup>	21 Sc 4s <sup>2</sup> 3d <sup>1</sup>	22 Ti 4s <sup>2</sup> 3d <sup>2</sup>	23 V 4s <sup>2</sup> 3d <sup>3</sup>	24 Cr 4s <sup>1</sup> 3d <sup>5</sup>	25 Mn 4s <sup>2</sup> 3d <sup>5</sup>	26 Fe 4s <sup>2</sup> 3d <sup>6</sup>	27 Co 4s <sup>2</sup> 3d <sup>7</sup>	28 Ni 4s <sup>2</sup> 3d <sup>8</sup>	29 Cu 4s <sup>1</sup> 3d <sup>10</sup>	30 Zn 4s <sup>2</sup> 3d <sup>10</sup>	31 Ga 4s <sup>2</sup> 4p <sup>1</sup>	32 Ge 4s <sup>2</sup> 4p <sup>2</sup>	33 As 4s <sup>2</sup> 4p <sup>3</sup>	34 Se 4s <sup>2</sup> 4p <sup>4</sup>	35 Br 4s <sup>2</sup> 4p <sup>5</sup>	36 Kr 4s <sup>2</sup> 4p <sup>6</sup>				
5	37 Rb 5s <sup>1</sup>	38 Sr 5s <sup>2</sup>	39 Y 5s <sup>2</sup> 4d <sup>1</sup>	40 Zr 5s <sup>2</sup> 4d <sup>2</sup>	41 Nb 5s <sup>1</sup> 4d <sup>4</sup>	42 Mo 5s <sup>1</sup> 4d <sup>5</sup>	43 Tc 5s <sup>2</sup> 4d <sup>5</sup>	44 Ru 5s <sup>1</sup> 4d <sup>7</sup>	45 Rh 5s <sup>1</sup> 4d <sup>8</sup>	46 Pd 4d <sup>10</sup>	47 Ag 5s <sup>1</sup> 4d <sup>10</sup>	48 Cd 5s <sup>2</sup> 4d <sup>10</sup>	49 In 5s <sup>2</sup> 5p <sup>1</sup>	50 Sn 5s <sup>2</sup> 5p <sup>2</sup>	51 Sb 5s <sup>2</sup> 5p <sup>3</sup>	52 Te 5s <sup>2</sup> 5p <sup>4</sup>	53 I 5s <sup>2</sup> 5p <sup>5</sup>	54 Xe 5s <sup>2</sup> 5p <sup>6</sup>				
6	55 Cs 6s <sup>1</sup>	56 Ba 6s <sup>2</sup>	57 La* 6s <sup>2</sup> 5d <sup>1</sup>	72 Hf 6s <sup>2</sup> 5d <sup>2</sup>	73 Ta 6s <sup>2</sup> 5d <sup>3</sup>	74 W 6s <sup>2</sup> 5d <sup>4</sup>	75 Re 6s <sup>2</sup> 5d <sup>5</sup>	76 Os 6s <sup>2</sup> 5d <sup>6</sup>	77 Ir 6s <sup>2</sup> 5d <sup>7</sup>	78 Pt 6s <sup>1</sup> 5d <sup>9</sup>	79 Au 6s <sup>1</sup> 5d <sup>10</sup>	80 Hg 6s <sup>2</sup> 5d <sup>10</sup>	81 Tl 6s <sup>2</sup> 6p <sup>1</sup>	82 Pb 6s <sup>2</sup> 6p <sup>2</sup>	83 Bi 6s <sup>2</sup> 6p <sup>3</sup>	84 Po 6s <sup>2</sup> 6p <sup>4</sup>	85 At 6s <sup>2</sup> 6p <sup>5</sup>	86 Rn 6s <sup>2</sup> 6p <sup>6</sup>				
7	87 Fr 7s <sup>1</sup>	88 Ra 7s <sup>2</sup>	89 Ac** 7s <sup>2</sup> 6d <sup>1</sup>	104 Rf 7s <sup>2</sup> 6d <sup>2</sup>	105 Db 7s <sup>2</sup> 6d <sup>3</sup>	106 Sg 7s <sup>2</sup> 6d <sup>4</sup>	107 Bh 7s <sup>2</sup> 6d <sup>5</sup>	108 Hs 7s <sup>2</sup> 6d <sup>6</sup>	109 Mt 7s <sup>2</sup> 6d <sup>7</sup>	110 Ds 7s <sup>2</sup> 6d <sup>8</sup>	111 Rg 7s <sup>2</sup> 6d <sup>9</sup>	112 Cn 7s <sup>2</sup> 6d <sup>10</sup>	113 Fl 7s <sup>2</sup> 7p <sup>1</sup>	114 Md 7s <sup>2</sup> 7p <sup>2</sup>	115 No 7s <sup>2</sup> 7p <sup>3</sup>	116 Lv 7s <sup>2</sup> 7p <sup>4</sup>	117 Lr 7s <sup>2</sup> 7p <sup>5</sup>	118 Lu 7s <sup>2</sup> 7p <sup>6</sup>				
*Lanthanides		58 Ce 6s <sup>2</sup> 4f <sup>1</sup> 5d <sup>1</sup>	59 Pr 6s <sup>2</sup> 4f <sup>3</sup>	60 Nd 6s <sup>2</sup> 4f <sup>4</sup>	61 Pm 6s <sup>2</sup> 4f <sup>5</sup>	62 Sm 6s <sup>2</sup> 4f <sup>6</sup>	63 Eu 6s <sup>2</sup> 4f <sup>7</sup>	64 Gd 6s <sup>2</sup> 4f <sup>7</sup> 5d <sup>1</sup>	65 Tb 6s <sup>2</sup> 4f <sup>9</sup>	66 Dy 6s <sup>2</sup> 4f <sup>10</sup>	67 Ho 6s <sup>2</sup> 4f <sup>11</sup>	68 Er 6s <sup>2</sup> 4f <sup>12</sup>	69 Tm 6s <sup>2</sup> 4f <sup>13</sup>	70 Yb 6s <sup>2</sup> 4f <sup>14</sup>	71 Lu 6s <sup>2</sup> 4f <sup>14</sup> 5d <sup>1</sup>							
6	**Actinides		90 Th 7s <sup>2</sup> 6d <sup>2</sup>	91 Pa 7s <sup>2</sup> 5f <sup>2</sup> 6d <sup>1</sup>	92 U 7s <sup>2</sup> f <sup>3</sup> 6d <sup>1</sup>	93 Np 7s <sup>2</sup> 5f <sup>4</sup> 6d <sup>1</sup>	94 Pu 7s <sup>2</sup> 5f <sup>6</sup>	95 Am 7s <sup>2</sup> 5f <sup>7</sup>	96 Cm 7s <sup>2</sup> 5f <sup>9</sup>	97 Bk 7s <sup>2</sup> 5f <sup>10</sup>	98 Cf 7s <sup>2</sup> 5f <sup>10</sup>	99 Es 7s <sup>2</sup> 5f <sup>11</sup>	100 Fm 7s <sup>2</sup> 5f <sup>12</sup>	101 Md 7s <sup>2</sup> 5f <sup>13</sup>	102 No 7s <sup>2</sup> 5f <sup>14</sup>	103 Lr 7s <sup>2</sup> 5f <sup>14</sup> 6d <sup>1</sup>						



# Chapter 7 & 8

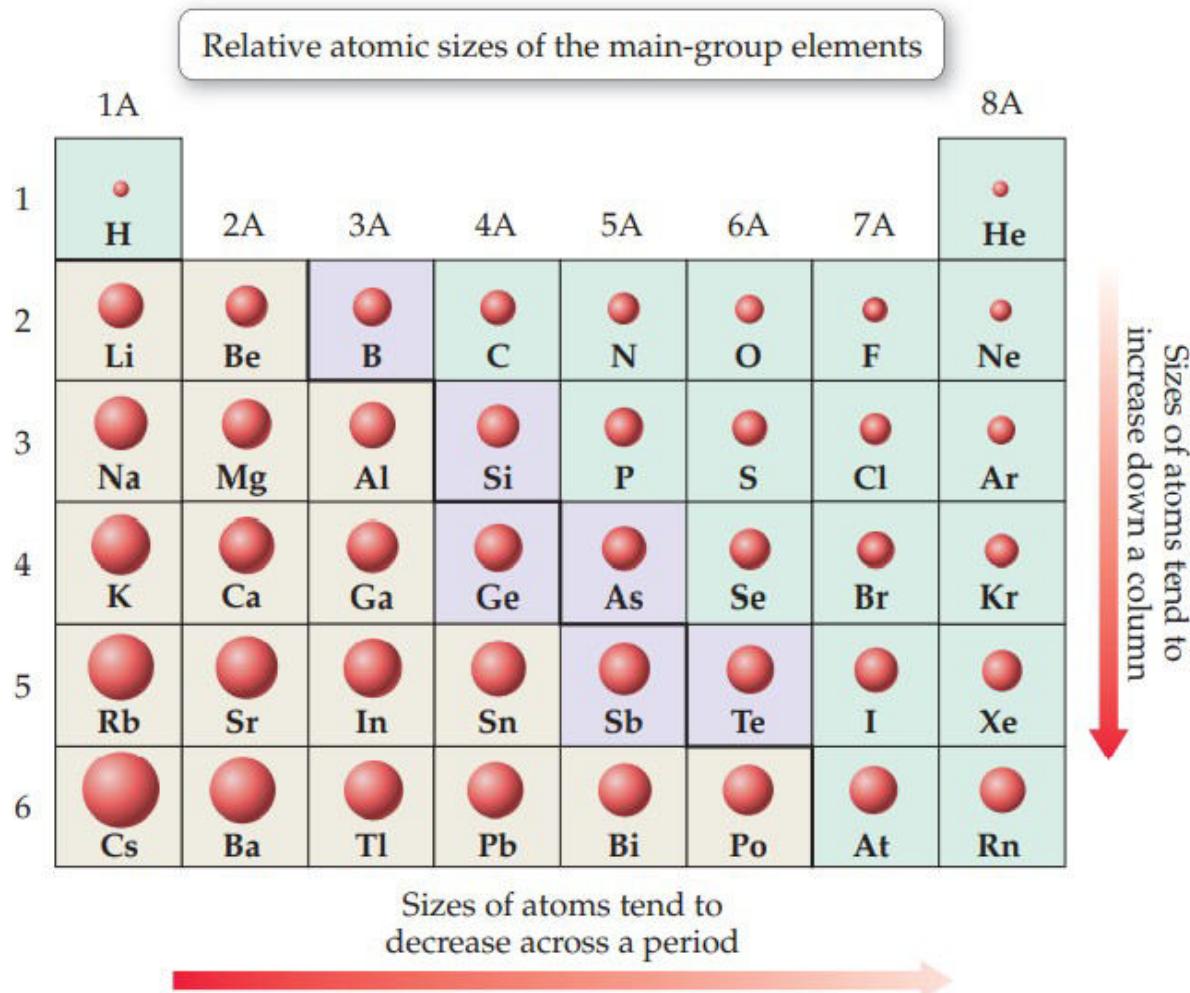
**Problem:** Using only the periodic table, arrange the following atoms in order of *increasing* atomic radius: P, Si, N.

**Solution:** From Figure we see that N and P are in the same group (Group 5A (15)). Therefore, the radius of N is smaller than that of P (atomic radius increases as we go down a group). Both Si and P are in the third period, and Si is to the left of P. Therefore, the radius of P is smaller than that of Si (atomic radius decreases as we move from left to right across a period). Thus, the order of increasing radius is: N < P < Si .





# Chapter 7 & 8



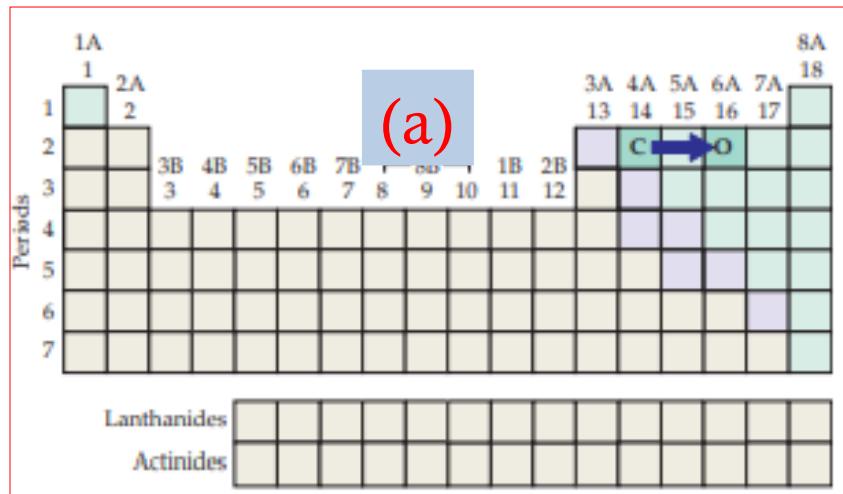
Periodic properties: atomic size



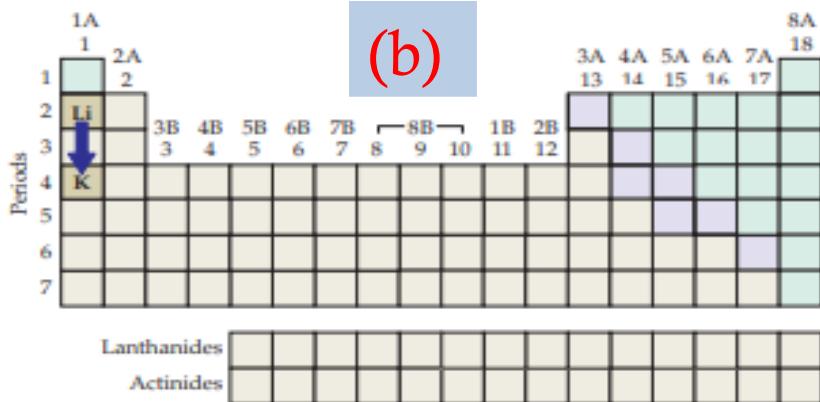
# Chapter 7 & 8

**Problem:** Choose the larger atom in each pair (atomic size). (a) C or O (b) Li or K (c) C or Al (d) Se or I.

**Solution:** (a) Carbon (C) atoms are larger than O atoms because, as you trace the path between C and O on the periodic table (see margin), you move to the right within the same period. Atomic size decreases as you go to the right.



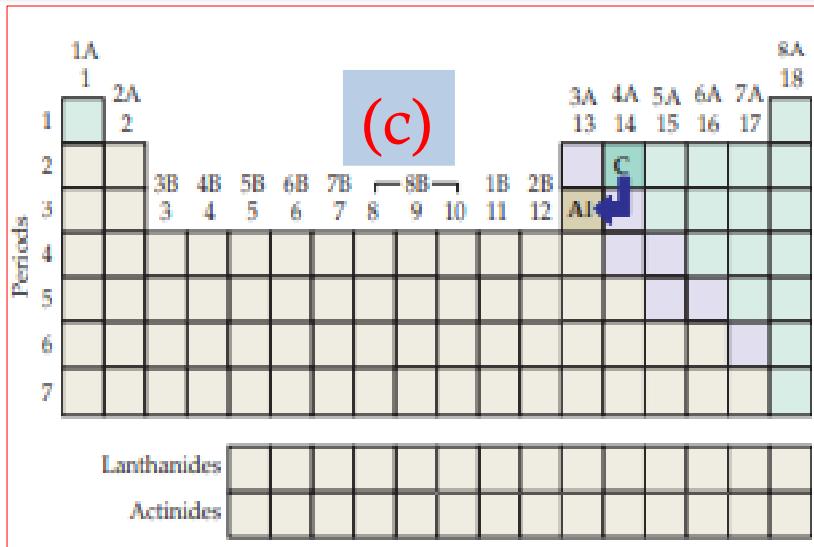
**Solution:** (b) Potassium atoms are larger than Li atoms because, as you trace the path between Li and K on the periodic table (see margin), you move down a column. Atomic size increases as you go down a column.



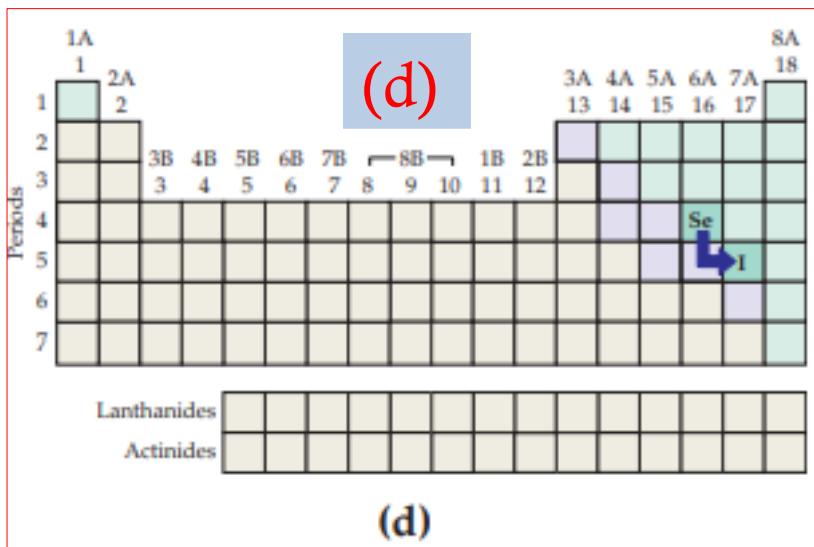


# Chapter 7 & 8

**Solution: (c)** Aluminum (Al) atoms are larger than C atoms because, as you trace the path between C and Al on the periodic table (see margin), you move down a column (atomic size increases) and then to the left across a period (atomic size increases). These effects add together for an overall increase.



**Solution: (d)** Based on periodic properties alone, you cannot tell which atom is larger because as you trace the path between Se and I (see margin) you go down a column (atomic size increases) and then to the right across a period (atomic size decreases). These effects tend to cancel one another.





# Chapter 7 & 8

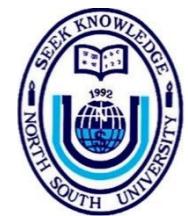
**Problem:** Using only the periodic table, rank the following in order of *decreasing* atomic size: (a) C, Mg, Sr (b) K, Ga, Ca (c) Br, Rb, Kr (d) Sr, Ca, Rb.

**Solution:** (a) Sr > Ca > Mg. These three elements are in Group (2), and size decreases up the group.

(b) K > Ca > Ga. These three elements are in Period 4, and size decreases across a period.

(c) Rb > Br > Kr. Rb is largest because it has one more energy level (Period 5) and is farthest to the left. Kr is smaller than Br because Kr is farther to the right in period 4.

(d) Rb > Sr > Ca. Ca is smallest because it has one fewer energy level. Sr is smaller than Rb because it is farther to the right.



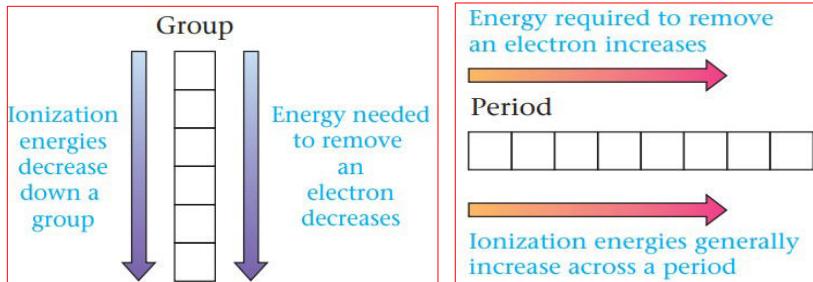
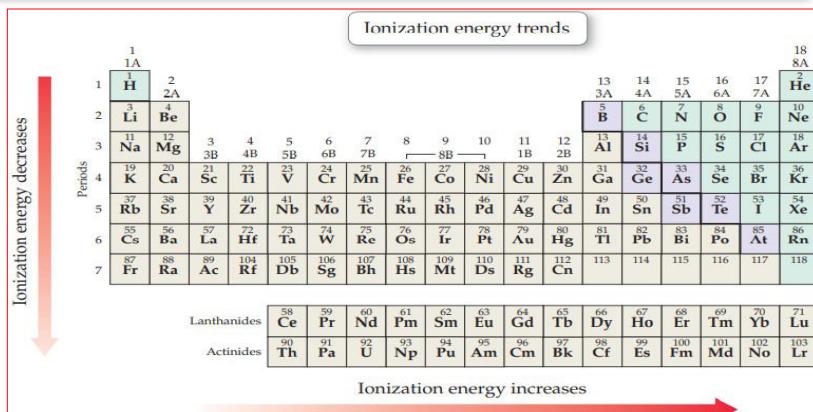
# Chapter 7 & 8

## • Trends in Ionization Energy

- Ionization energy (IE) is the energy required to remove 1 mol of electrons from 1 mol of gaseous atoms or ions.

- Metals have a *low IE* and they form **+ve ions** or **cations**
- Non-metals have a *high IE* and they form **-ve ions** or **anions**.

- Ionization energy tends to *increase* across a period and *decrease* down a group.



Metals at the bottom of a group lose electrons more easily than those at the top.

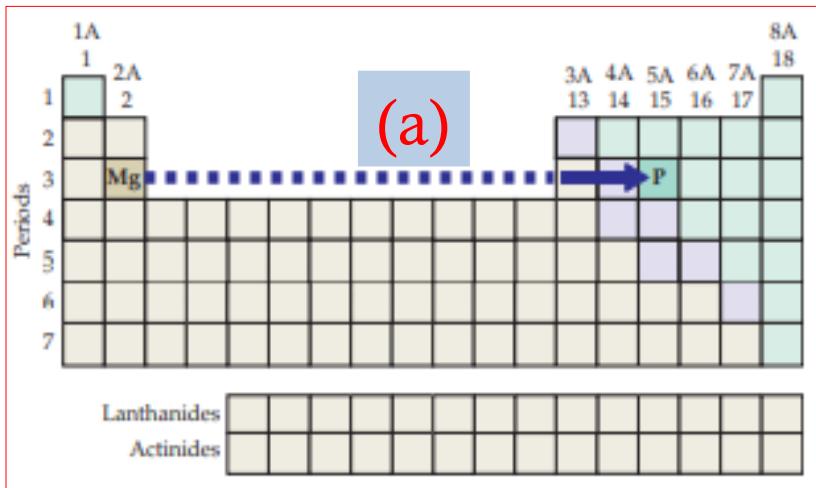
Metals appear on the left side of the periodic table and nonmetals appear on the right. ionization energies tend to increase from left to right across a given period on the periodic table.



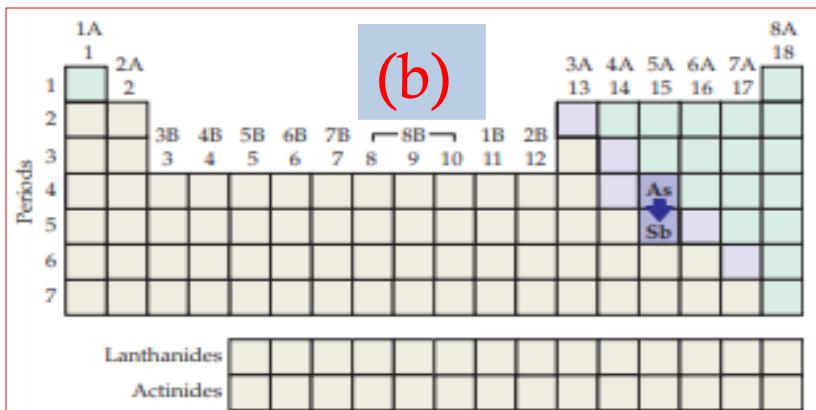
# Chapter 7 & 8

**Problem:** Choose the element with the higher ionization energy from each pair. (a) Mg or P (b) As or Sb (c) N or Si (d) O or Cl.

**Solution:** (a) P has a higher ionization than Mg because, as you trace the path between Mg and P on the periodic table (see margin), you move to the right within the same period. Ionization energy increases as you go to the right.



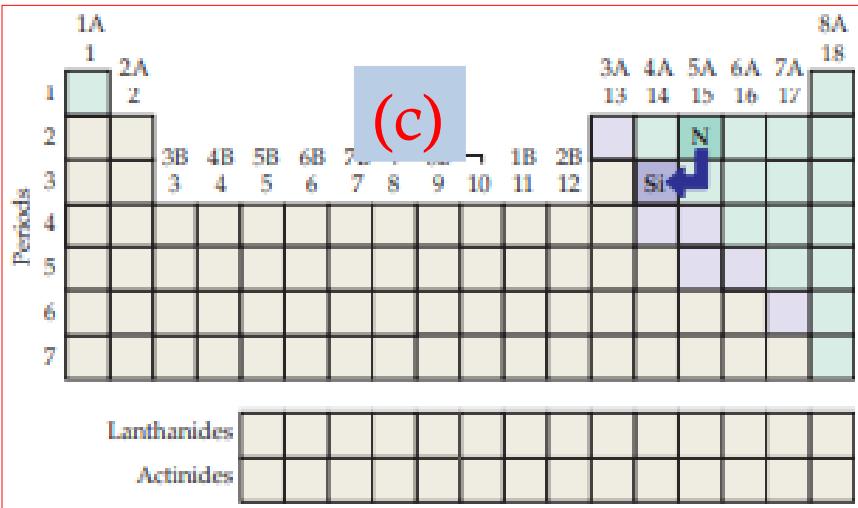
**Solution:** (b) As has a higher ionization energy than Sb because, as you trace the path between As and Sb on the periodic table (see margin), you move down a column. Ionization energy decreases as you go down a column.



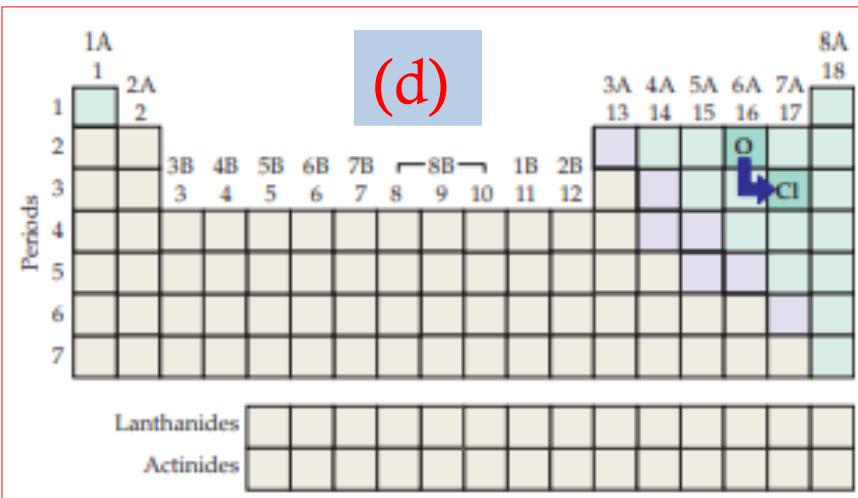


# Chapter 7 & 8

**Solution: (c)** N has a higher ionization energy than Si because, as you trace the path between N and Si on the periodic table (see margin), you move down a column (ionization energy decreases) and then to the left across a period (ionization energy decreases). These effects sum together for an overall decrease.



**Solution: (d)** Based on periodic properties alone, you cannot tell which has a higher ionization energy because, as you trace the path between O and Cl (see margin), you go down a column (ionization energy decreases) and then to the right across a period (ionization energy increases). These effects tend to cancel.

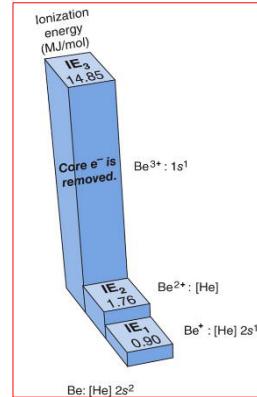
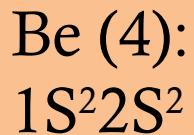




# Chapter 7 & 8

**Problem:** Using the periodic table, rank the following in order of decreasing  $\text{IE}_1$ : (a) Kr, He, Ar (b) Sb, Te, Sn (c) K, Ca, Rb (d) I, Xe, Cs.

The first three Ionization Energies of Beryllium



IE energy drastically increasing due to decreasing outer shell electrons

**Solution:** (a)  $\text{He} > \text{Ar} > \text{Kr}$ . These are in Group 8A (18), and  $\text{IE}_1$  decreases down a group.

(b)  $\text{Te} > \text{Sb} > \text{Sn}$ . These are in Period 5, and  $\text{IE}_1$  increases across a period.

(c)  $\text{Ca} > \text{K} > \text{Rb}$ .  $\text{IE}_1$  of K is larger than  $\text{IE}_1$  of Rb because K is higher in Group 1A (1).  $\text{IE}_1$  of Ca is larger than  $\text{IE}_1$  of K because Ca is farther to the right in period 4.

(d)  $\text{Xe} > \text{I} > \text{Cs}$ .  $\text{IE}_1$  of I is smaller than  $\text{IE}_1$  of Xe because I is farther to the left of I is larger than  $\text{IE}_1$  of Cs because I is farther to the right and in the previous period.

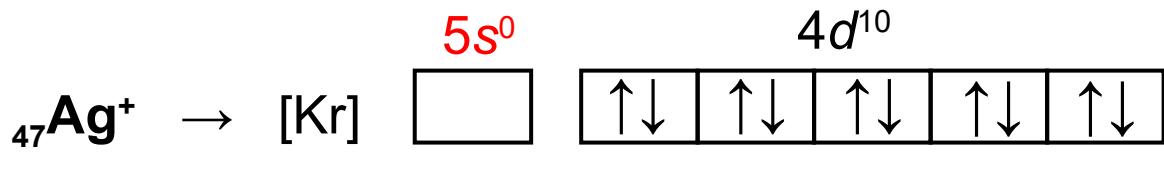
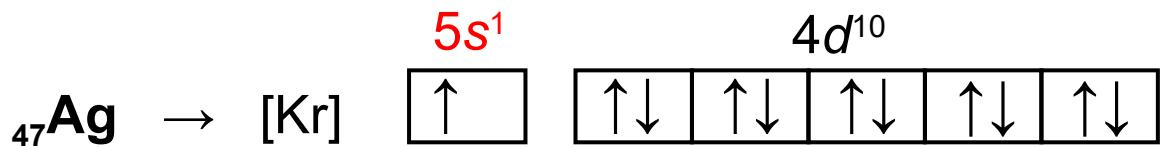


# Chapter 7 & 8

## •Magnetic Properties of Elements

### Paramagnetic vs Diamagnetic

- A species with any **unpaired** electrons is **paramagnetic**, i.e. it is attracted by a magnetic field.
- A species with all its electrons **paired** is **diamagnetic**, i.e. it is not attracted (and is slightly repelled) by a magnetic field.
- Consistent with the **First-In First-Out Principle**, Ag is **paramagnetic** whereas Ag<sup>+</sup> turns out to be **diamagnetic**.

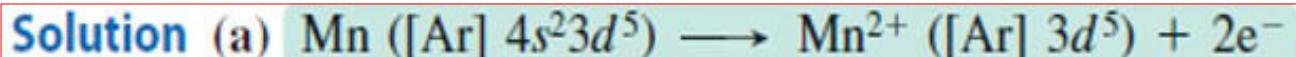
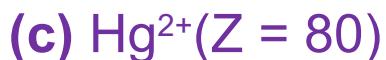
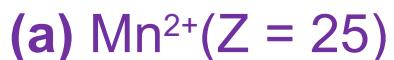




# Chapter 7 & 8

## Sample Problem:

Use condensed electron configurations for each transition metal ion, and predict whether the ion is paramagnetic or diamagnetic.



There are five unpaired  $e^-$ , so  $\text{Mn}^{2+}$  is paramagnetic.



There are three unpaired  $e^-$ , so  $\text{Cr}^{3+}$  is paramagnetic.



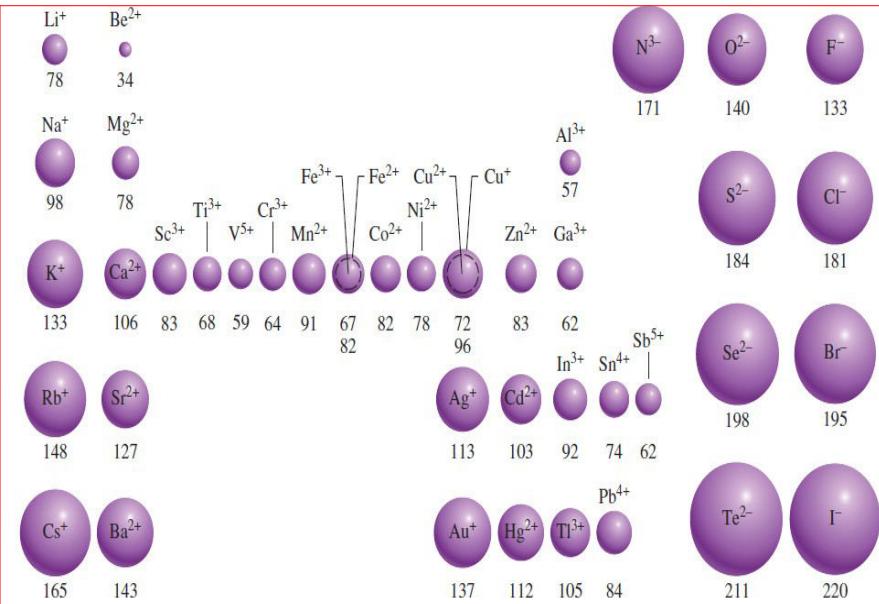
The  $4f$  and  $5d$  sublevels are filled, so there are no unpaired  $e^-$ :  $\text{Hg}^{2+}$  is not paramagnetic.

# Chapter 7 & 8

## • Trends in Ionic Size

- Cations are *smaller* than their parent atoms while anions are *larger*.
- e.g.  $\text{Na}^+$  ion has a smaller radius than Na atom. On the other hand,  $\text{Cl}^-$  ion has a larger radius than Cl atom.

- Cation size *decreases* as the net positive charge *increases*.
- e.g.  $\text{Mg}^{2+}$  ion has a smaller radius than  $\text{Na}^+$  ion.





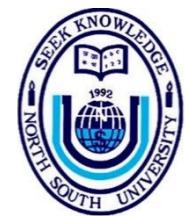
# Chapter 7 & 8

For each of the following pairs, indicate which one of the two species is larger: (a)  $\text{N}^{3-}$  or  $\text{F}^-$ ; (b)  $\text{Mg}^{2+}$  or  $\text{Ca}^{2+}$ ; (c)  $\text{Fe}^{2+}$  or  $\text{Fe}^{3+}$ .

**Solution** (a)  $\text{N}^{3-}$  and  $\text{F}^-$  are isoelectronic anions, both containing 10 electrons.

Because  $\text{N}^{3-}$  has only seven protons and  $\text{F}^-$  has nine, the smaller attraction exerted by the nucleus on the electrons results in a larger  $\text{N}^{3-}$  ion.

- (b) Both Mg and Ca belong to Group 2A (the alkaline earth metals). Thus,  $\text{Ca}^{2+}$  ion is larger than  $\text{Mg}^{2+}$  because Ca's valence electrons are in a larger shell ( $n = 4$ ) than are Mg's ( $n = 3$ ).
- (c) Both ions have the same nuclear charge, but  $\text{Fe}^{2+}$  has one more electron (24 electrons compared to 23 electrons for  $\text{Fe}^{3+}$ ) and hence greater electron-electron repulsion. The radius of  $\text{Fe}^{2+}$  is larger.



# Chapter 7 & 8

## • Trends in Metallic

- Metals tend to lose electrons in their chemical reactions, while nonmetals tend to gain electrons.
- Move across a period, or row, to the right in the periodic table, metallic character decreases.
- Move down a column, or family, in the periodic table, metallic character increases.

	1 2	13 14 15 16 17 18
H <sup>+</sup>	H	He
He	Li Be	B C N O F Ne
Ne	Na Mg	Al Si P S Cl Ar
Ar	K Ca	Ga Ge As Se Br Kr
Kr	Rb Sr	In Sn Sb Te I Xe

Metals Tend to Lose Electrons

Metallic character increases  
↓

Periods	1 IA	2 IA	3 IA	4 IA	5 IA	6 IA	7 IA	8 IA	9 IA	10 IA	11 IA	12 IA	13 IA	14 IA	15 IA	16 IA	17 IA	18 IA
1	H	Be																He
2	Li	Mg																
3	Na	Mg	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
4	K	Ca																
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Pt	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac	Db	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	113	114	115	116	117	118

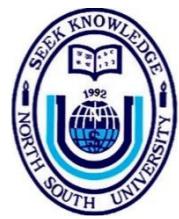
Metallic character trends

Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Metallic character decreases  
→

1 2	13 14 15 16 17 18	He
H		
Li Be	B C N O F	Ne
Na Mg	Al Si P S Cl	Ar
K Ca	Ga Ge As Se Br	Kr
Rb Sr	In Sn Sb Te I	Xe

Nonmetals Tend to Gain Electrons

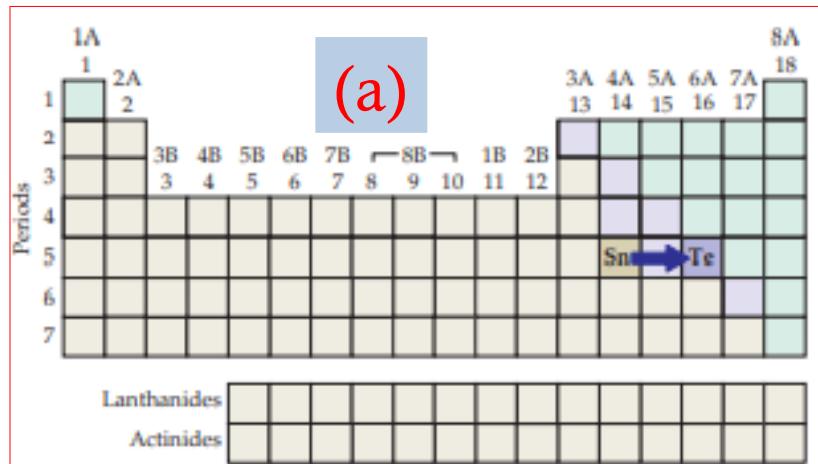


# Chapter 7 & 8

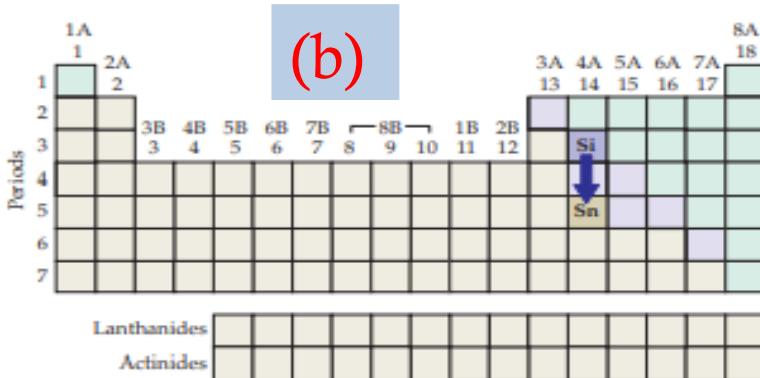
**Problem:** Choose the more metallic element from each pair.

- (a) Sn or Te (b) Si or Sn (c) Br or Te (d) Se or I.

**Solution:** (a) Sn is more metallic than Te because, as you trace the path between Sn and Te on the periodic table (see margin), you move to the right within the same period. Metallic character decreases as you go to the right.



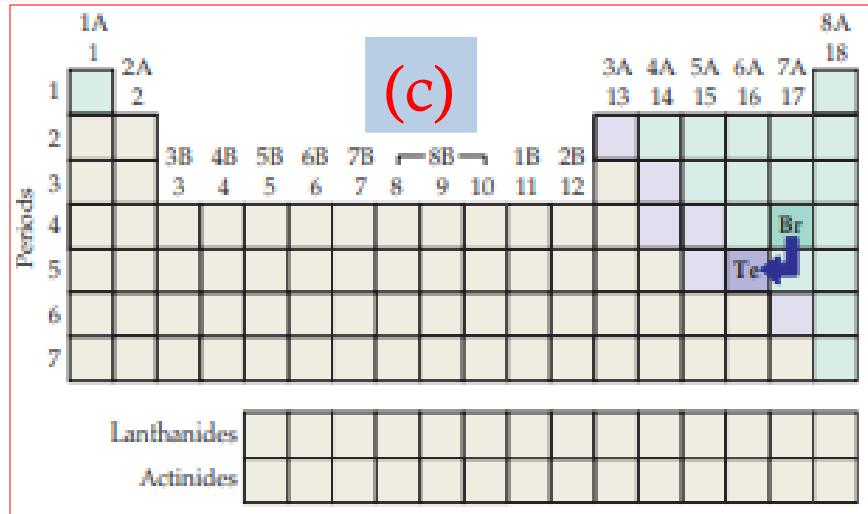
**Solution:** (b) Sn is more metallic than Si because, as you trace the path between Si and Sn on the periodic table (see margin), you move down a column. Metallic character increases as you go down a column.



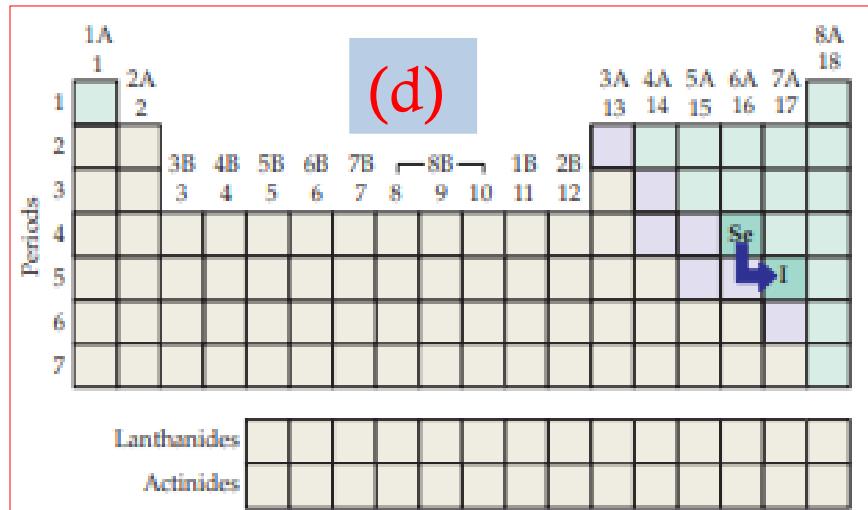


# Chapter 7 & 8

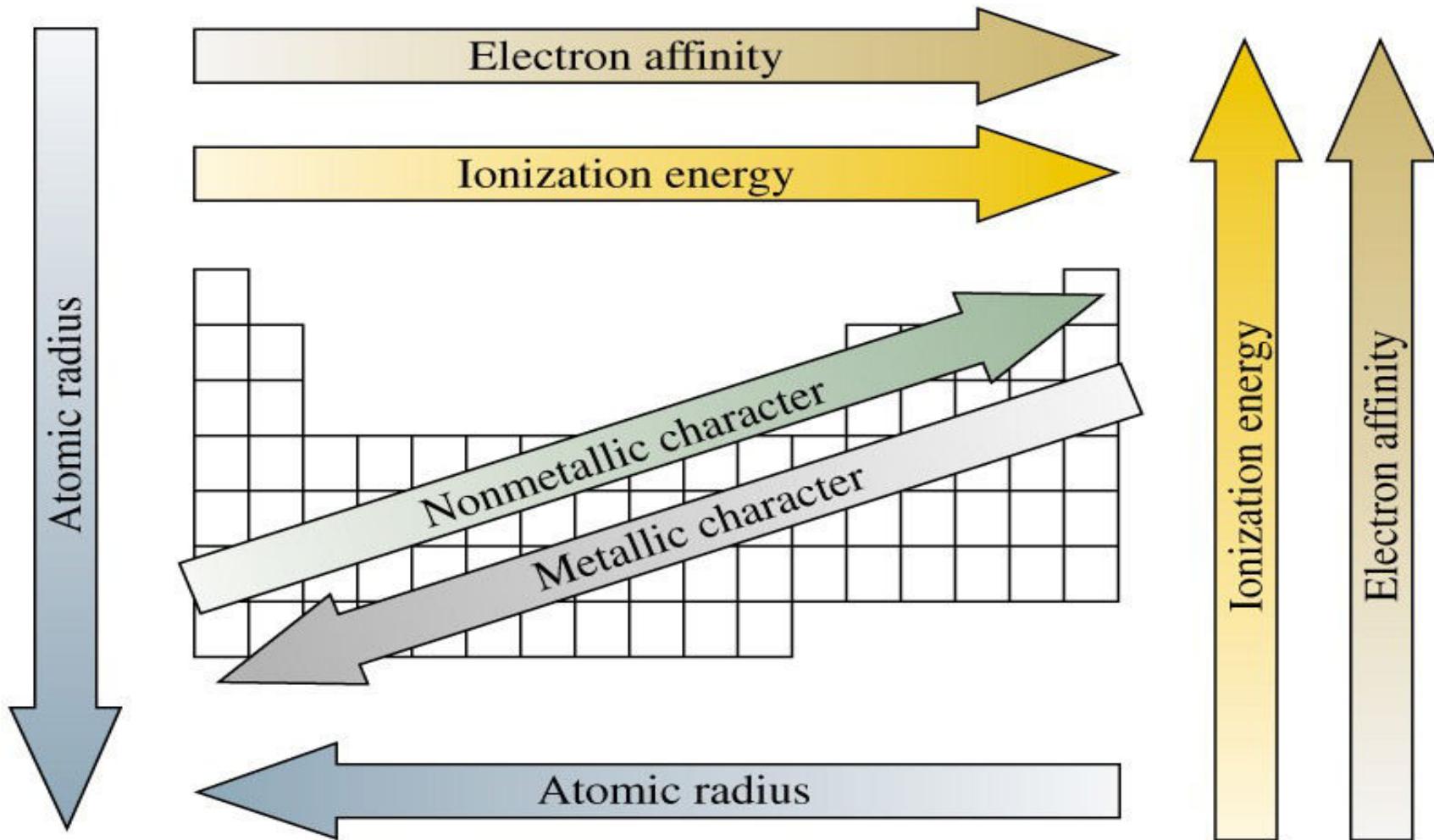
**Solution:** (c) Te is more metallic than Br because, as you trace the path between Br and Te on the periodic table (see margin), you move down a column (metallic character increases) and then to the left across a period (metallic character increases). These effects add together for an overall increase.



**Solution:** (d) Based on periodic properties alone, you cannot tell which is more metallic because as you trace the path between Se and I (see margin), you go down a column (metallic character increases) and then to the right across a period (metallic character decreases). These effects tend to cancel.



# Chapter 7 & 8



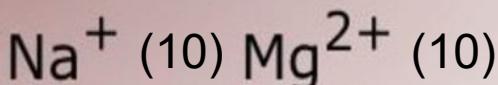
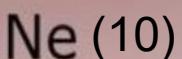
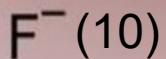
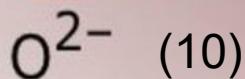
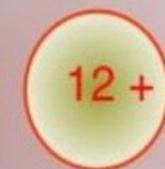
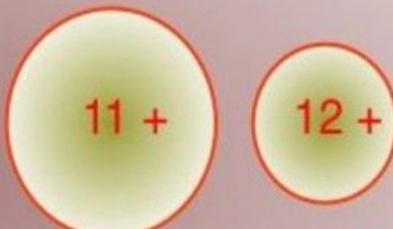
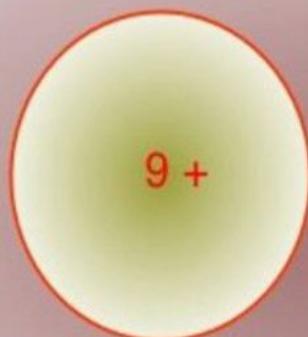
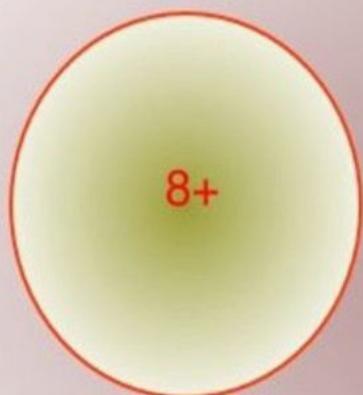
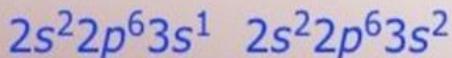
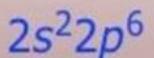
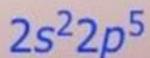
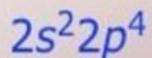


# Chapter 7 & 8

## Isoelectronic

Ions with the same electronic configuration as a noble gas are said to be isoelectronic with a noble gas.

Electron configurations of the neutral atoms:



**Isoelectronic series:** all these species have ten electrons:



# Chapter 7 & 8

## • $\text{K}^+$ Stability

- Potassium atom (19):  $1s^2 2s^2 2p^6 3s^2 3p^6 \textcolor{blue}{4s}^1$
- $\text{K}^+$  ion:  $1s^2 2s^2 2p^6 3s^2 3p^6$
- Ar atom:  $1s^2 2s^2 2p^6 3s^2 3p^6$

$\text{K}^+$  and Ar are isoelectronic, that is, they have identical electron configurations



# Chapter 7 & 8

## • $\text{Ca}^{2+}$ Stability

- Calcium atom (20):  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
- $\text{Ca}^{2+}$  ion:  $1s^2 2s^2 2p^6 3s^2 3p^6$
- Ar atom:  $1s^2 2s^2 2p^6 3s^2 3p^6$

$\text{Ca}^{2+}$  and Ar are **isoelectronic**, that is, they have identical electron configurations



# Chapter 7 & 8

What ion charge do you think would be particularly stable for a Group 3A element like aluminum?

1. -3
2. 0
3. +1
4. +3

Aluminum forms a **stable**  $\text{Al}^{3+}$  cation which is isoelectronic with the preceding noble gas Ne.





# Chapter 7 & 8

## • Ions of Nonmetals

- Chlorine has an electron configuration one short of a noble gas



- Picking up one extra electron completes an argon configuration





# Chapter 7 & 8

## •Ions of Nonmetals

**Effective nuclear charge ( $Z_{\text{eff}}$ )** is the “positive charge” felt by an electron.

$$Z_{\text{eff}} = Z - \sigma$$

$0 < \sigma < Z$  ( $\sigma$  = shielding constant)

	<u><math>Z</math></u>	<u>Core</u>	<u><math>Z_{\text{eff}}</math></u>	<u>Radius (pm)</u>
$Z_{\text{eff}} \approx Z - \text{number of inner or core electrons}$				

Na	11	10	1	186	Na (11): $1s^2 2s^2 2p^6 3s^1$
Mg	12	10	2	160	
Al	13	10	3	143	
Si	14	10	4	132	

Thank You for Listening  
ດីជានុក សិរី ទៅអ្នក



Any Question  
នូវ សោចនា