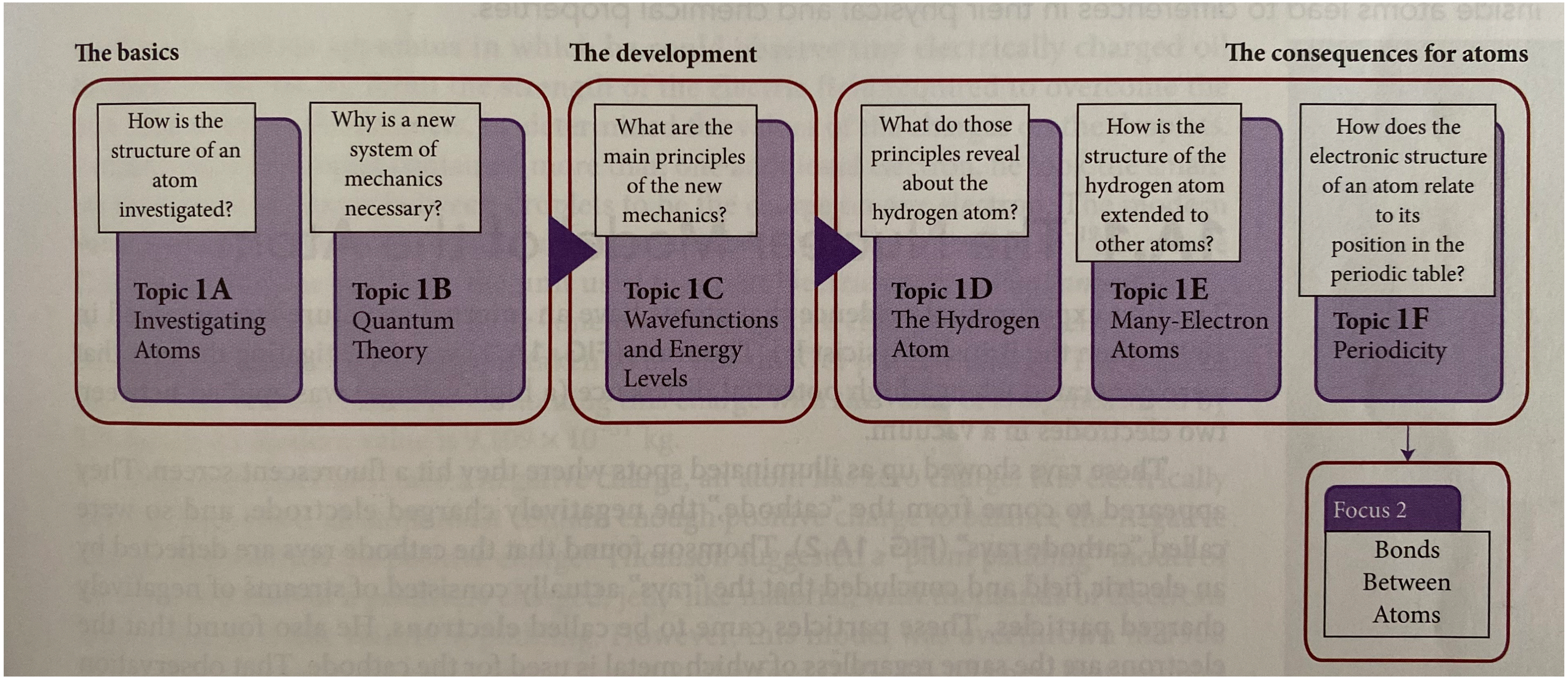
CH-110 Advanced General Chemistry I

Prof. A. Steinauer

angela.steinauer@epfl.ch

Overview Chapter 1 (Focus 1: Atoms)



The basics

How is the structure of an atom investigated?

Topic 1A Investigating Atoms

Why is a new system of mechanics necessary?

Topic 1B Quantum Theory

The development

What are the main principles of the new mechanics?

Topic 1C Wavefunctions and Energy Levels

The consequences for atoms

What do those principles reveal about the hydrogen atom?

Topic 1D The Hydrogen Atom

How is the structure of the hydrogen atom extended to other atoms?

Topic 1E Many-Electron Atoms

How does the electronic structure of an atom relate to its position in the periodic table?

Topic 1F Periodicity

Focus 2

Bonds Between Atoms

Topic 1D.1 Energy levels
Topic 1D.2 Atomic orbitals
Topic 1D.3 Quantum numbers, shells, and subshells
WE CONTINUE HERE: Topic 1D.4 The shapes of orbitals
Topic 1D.5 Electron spin
Topic 1D.6 The electronic structure of hydrogen: a summary

WHY DO YOU NEED TO KNOW THIS MATERIAL?

- The hydrogen atom is the **simplest atom of all** and is used to discuss the **structures of all atoms**.
- It is therefore **central** to many explanations in chemistry.

WHAT DO YOU NEED TO KNOW ALREADY?

- Features of spectrum of atomic hydrogen (Topic 1A)
- Concepts of wavefunction and energy level in quantum mechanics (Topic 1C)

1D.4 The shapes of orbitals

Boundary surface representation

- Encloses most of the electron cloud
- Easier to draw, but atoms really have fuzzy edges
- Still useful: shows where the electron is most likely found.

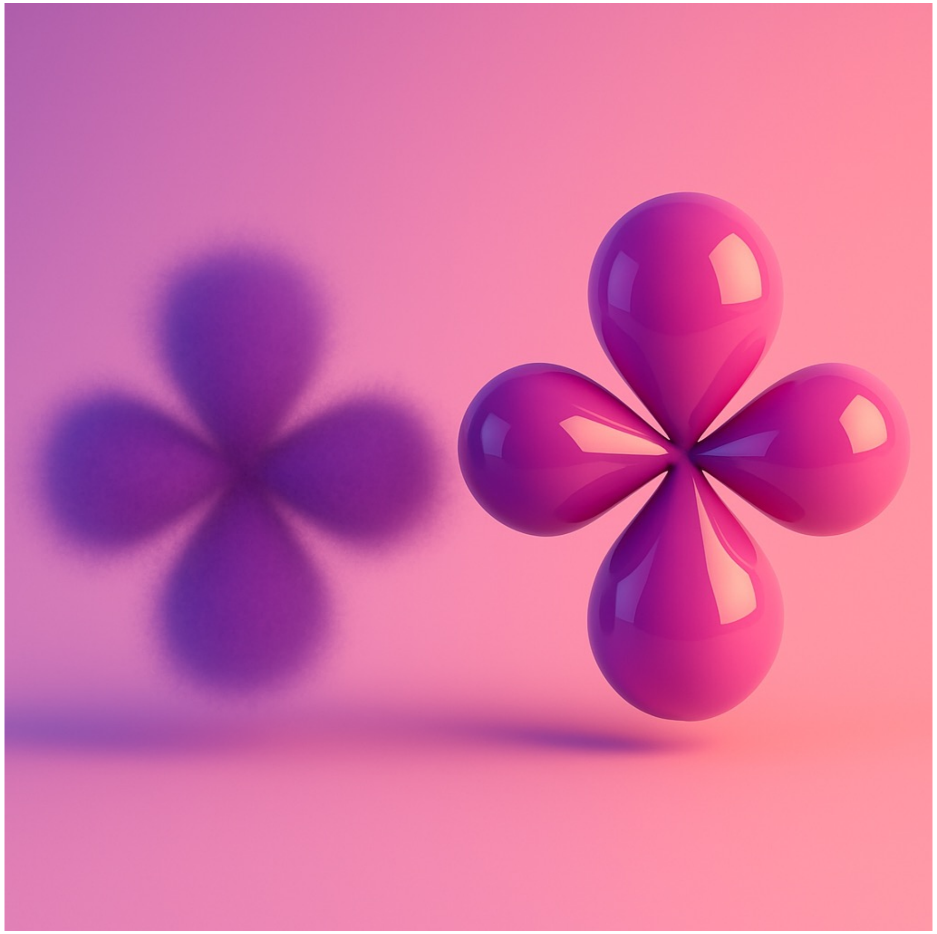


Image source: ChatGPT (2025).
Fuzzy Real vs. Boundary Surface.

1D.4 The shapes of orbitals

Boundary surface of the 1s-orbital

- Keep in mind:

  The probability density inside the boundary surface is not uniform.

- An s-orbital has a spherical boundary surface because the electron cloud is spherical.

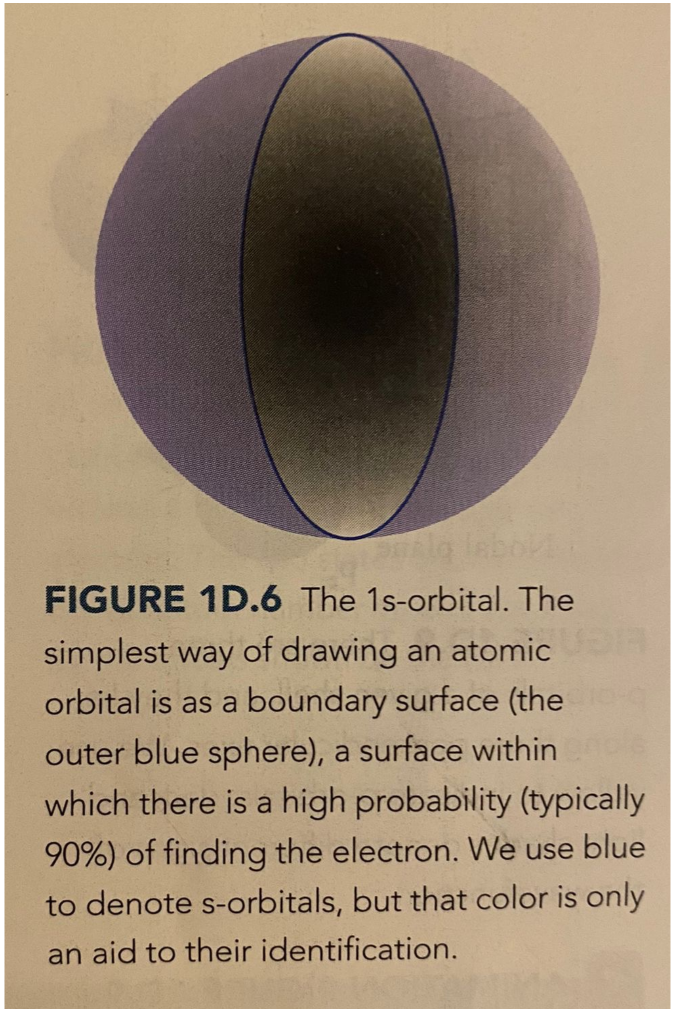


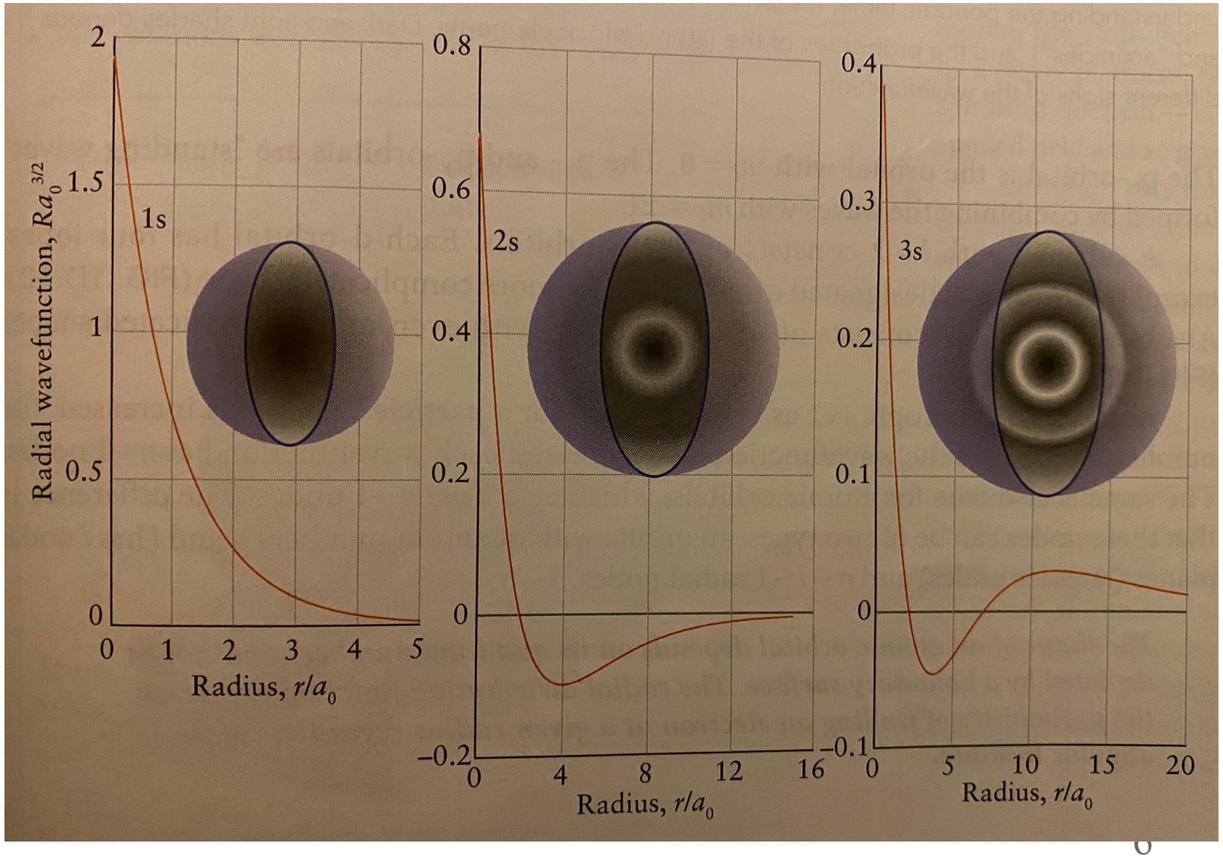
FIGURE 1D.6 The 1s-orbital. The simplest way of drawing an atomic orbital is as a boundary surface (the outer blue sphere), a surface within which there is a high probability (typically 90%) of finding the electron. We use blue to denote s-orbitals, but that color is only an aid to their identification.

1D.3 Quantum numbers, shells, and subshells

Boundary surfaces of higher-order s-orbitals

- The s-orbitals of higher energy have spherical boundary surfaces of greater diameter
- They also have a more complicated radial variation, with **radial nodes** (radii at which the wavefunction passes through zero).

Figure 1D.7



1D.3 Quantum numbers, shells, and subshells

Boundary surfaces of higher-order s-orbitals

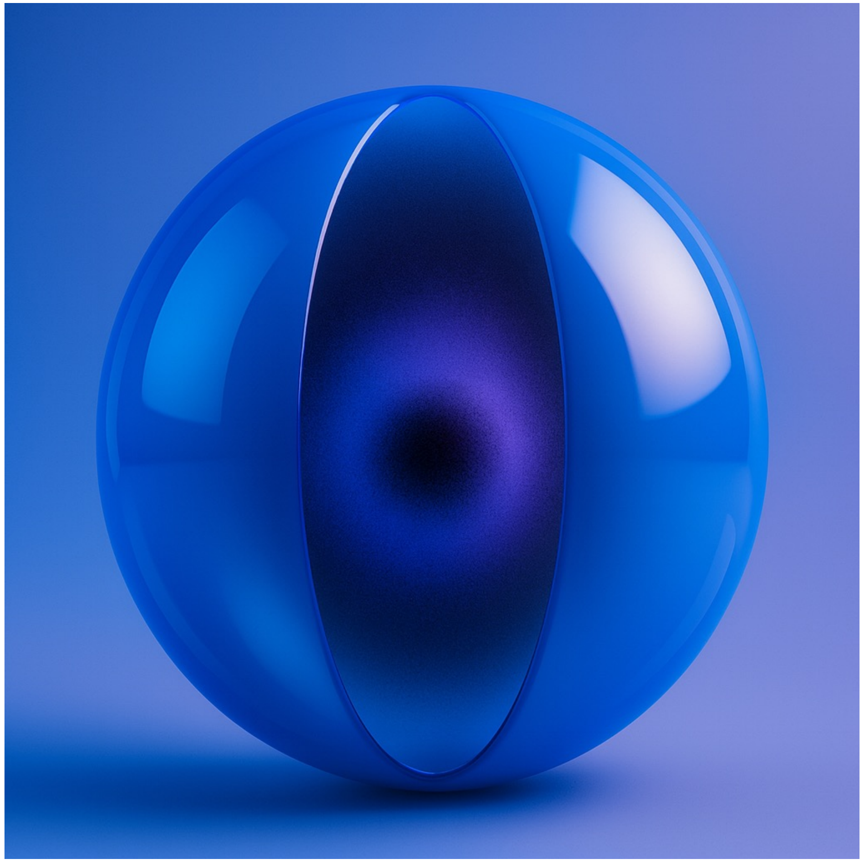


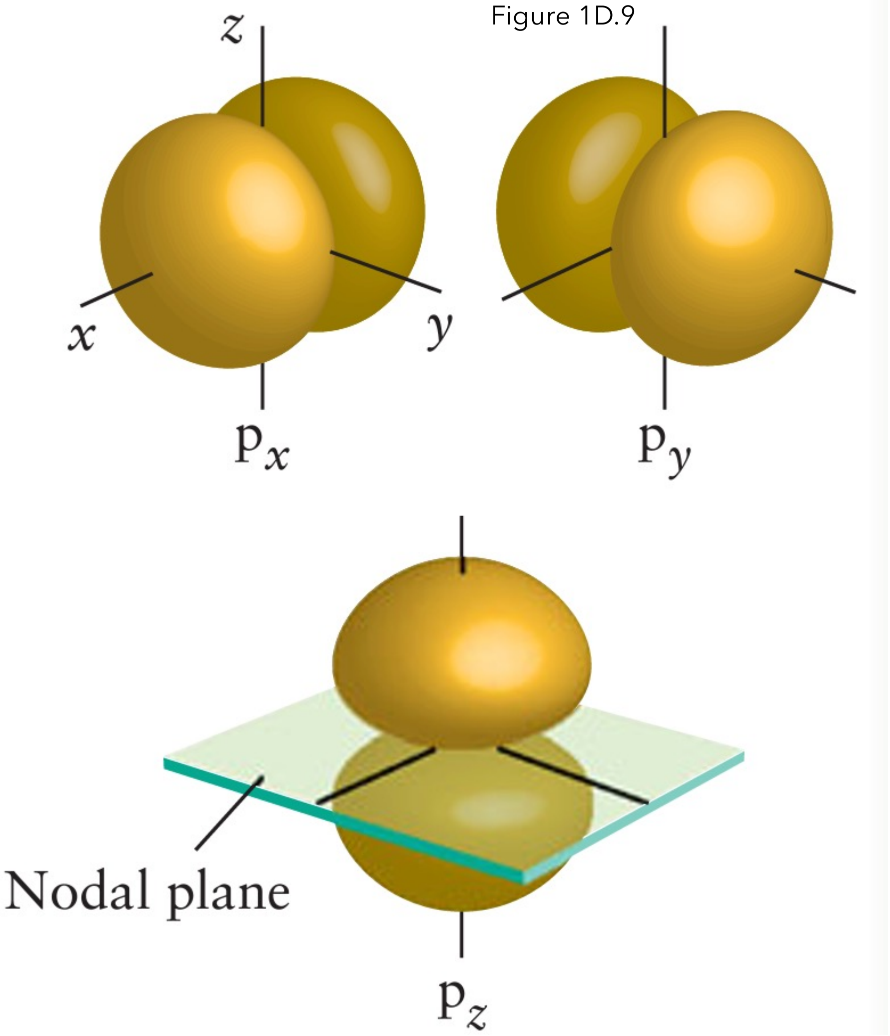
Image source: ChatGPT (2025).
A Glimpse inside a 2s Orbital.

1D.4 The shapes of orbitals

p-orbitals

- Shape: two lobes with opposite signs of the wavefunction (+/–)
- Separated by a **nodal plane** through the nucleus where $\psi = 0$ (electron never at the nucleus)
- Example: $2\mathrm{p_z} \propto \cos(\theta)$, which changes sign across the nodal plane
- Three p-orbitals per subshell: $\mathrm{p_x, p_y, p_z}$
- Quantum numbers: $m_l = -1, 0, +1$
- Electrons in p-orbitals have nonzero angular momentum

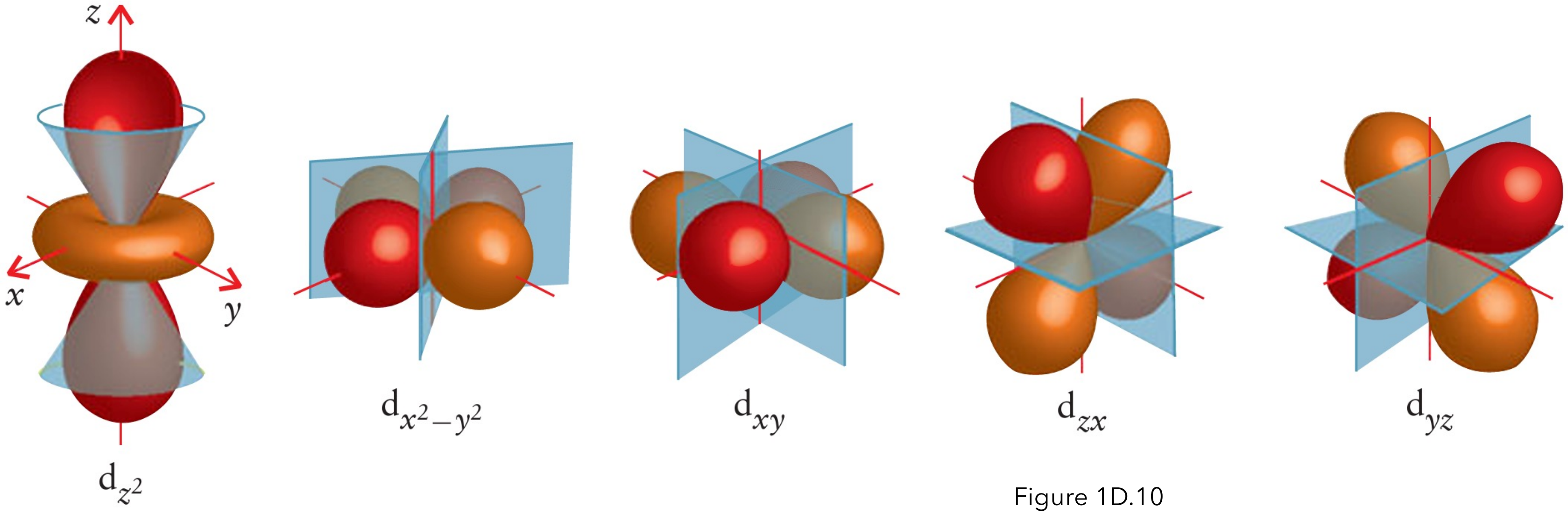
Figure 1D.9



1D.4 The shapes of orbitals

d-orbitals

- Subshell $l = 2$ consists of **five d-orbitals**
- Each d-orbital has four lobes, except d_{z^2}



d_{z^2} $d_{x^2-y^2}$ d_{xy} d_{zx} d_{yz}

Figure 1D.10

1D.4 The shapes of orbitals

f-orbitals

- Subshell $l = 3$ consists of **seven f-orbitals**.
- Very **complex** appearance.
- Detailed form will not be discussed again in this course.
- Their existence is important for **understanding the periodic table** (lanthanoids and actinoids).

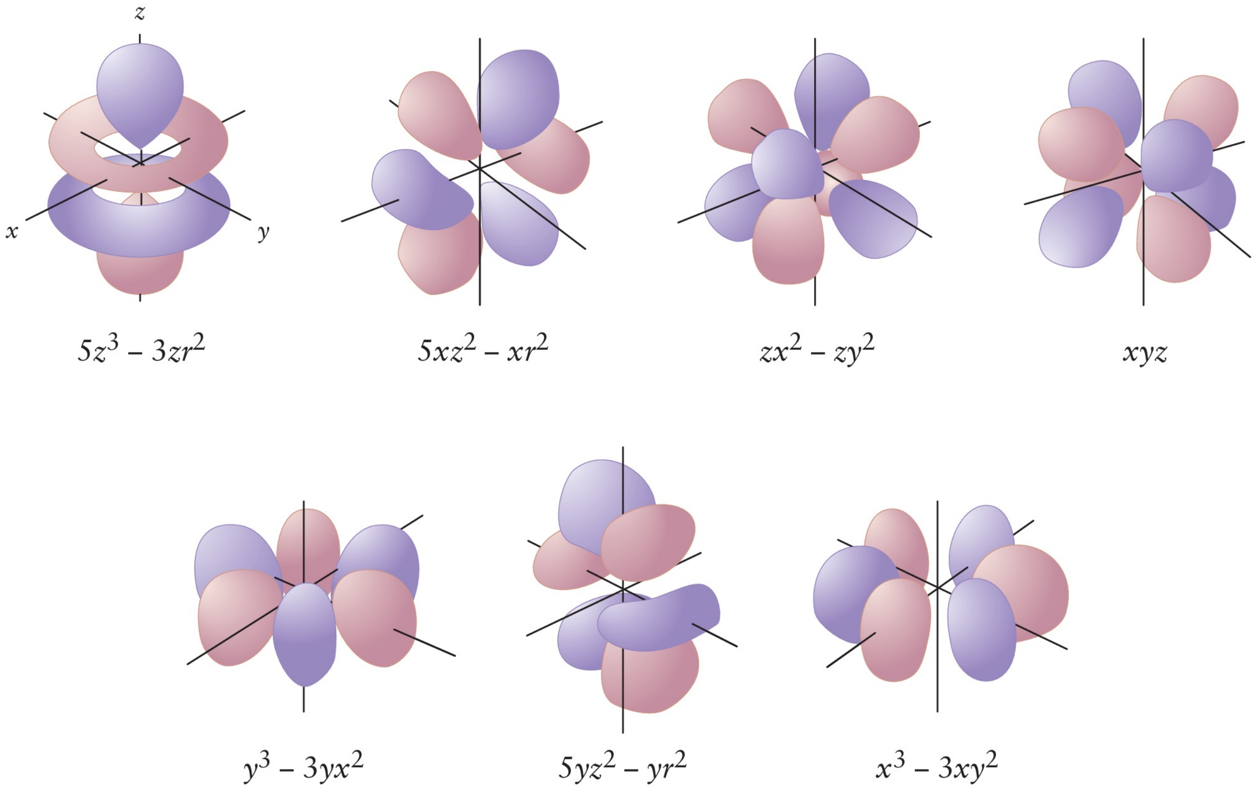


Figure 1D.11

1D.4 The shapes of orbitals

Summary

The shape of an atomic orbital depends on its quantum numbers and can be depicted by a boundary surface. The radial distribution function expresses the probability of finding an electron at a given radius regardless of its angular momentum.

Electron Spin

Topic 1D.5

1D.5 Electron spin

Spin ↑ and ↓

Think of an electron as being able to spin counterclockwise (the ↑ state) and clockwise (the ↓ state) at exactly the same rate.

These two spins are distinguished by a **fourth quantum number**, the spin magnetic quantum number, m_s.

This quantum number can have only one of two values: $+\frac{1}{2}$ (↑) and $-\frac{1}{2}$ (↓).

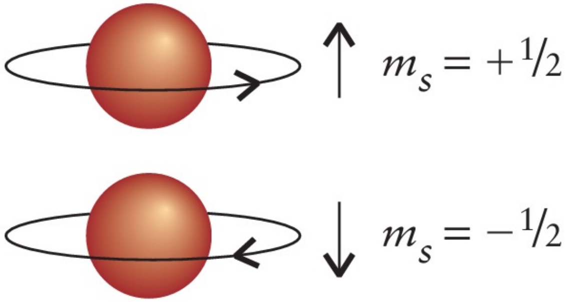


Figure 1D.12

1D.5 Electron spin

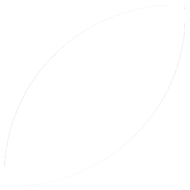
Summary

An electron has the property of spin;

the spin is described by the quantum number $m_s = \pm\frac{1}{2}$.

The Electronic Structure of Hydrogen: A Summary

Topic 1D.6



1D.6 The electronic structure of hydrogen: a summary

Orbital summary

1D.6 The electronic structure of hydrogen: a summary

1) In the ground state of hydrogen:

$$n = 1, l = 0, m_l = 0, m_s = \pm\frac{1}{2}$$

Both values of m_s are possible, spin orientation does not affect energy.

2) When an atom acquires enough energy (by absorbing a photon) for its electron to reach n = 2:

It can occupy any of the four orbitals in that shell: one 2s and three 2p orbitals (in hydrogen, they all have the same energy)

1D.6 The electronic structure of hydrogen: a summary

3) Atom acquires even more energy:

Electron can move to *n* = 3 shell

Atom is now even larger

Nine orbitals available (3s, 3p, 3d)

4) More energy still:

Electron can move to *n* = 4 shell with 16 available orbitals

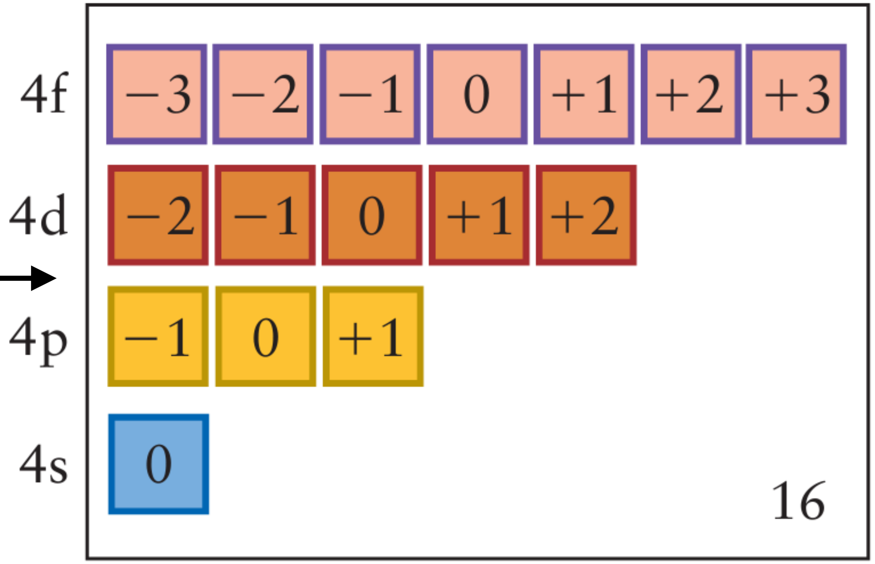


Figure 1D.13

1D.6 The electronic structure of hydrogen: a summary

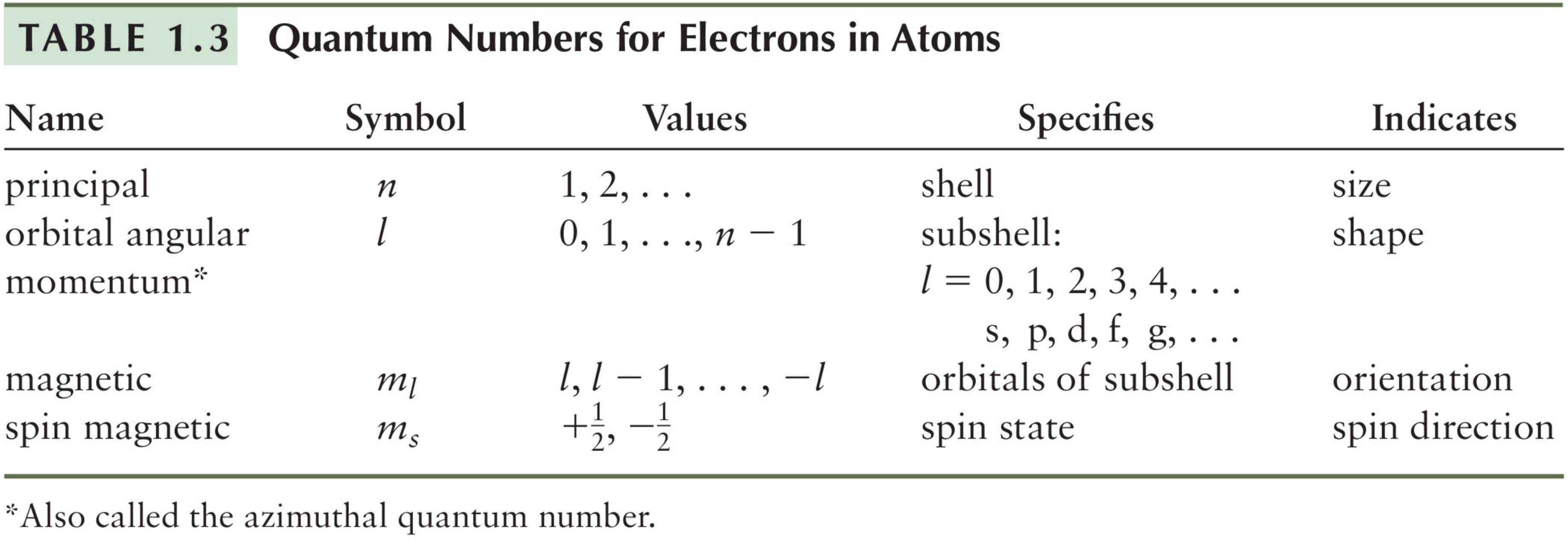


TABLE 1.3 Quantum Numbers for Electrons in Atoms

Name	Symbol	Values	Specifies	Indicates
principal	n	$1, 2, \ldots$	shell	size
orbital angular momentum*	l	$0, 1, \ldots, n-1$	subshell: $l = 0, 1, 2, 3, 4, \ldots$ s, p, d, f, g, . . .	shape
magnetic	m_l	$l, l-1, \ldots, -l$	orbitals of subshell	orientation
spin magnetic	m_s	$+\frac{1}{2}, -\frac{1}{2}$	spin state	spin direction

*Also called the azimuthal quantum number.

1D.6 The electronic structure of hydrogen: a summary

Summary

The state of an electron in a hydrogen atom is defined by the four quantum numbers n, l, m_l and m_s; as the value of n increases, the size of the atom increases.

The skills you have mastered are the ability to

- ❑ Assess the relative probability of finding an electron at a given distance from the nucleus of an atom.
- ❑ Name and explain the relation of each of the four quantum numbers to the properties and relative energies of atomic orbitals.
- ❑ Describe the properties of electron spin.
- ❑ Describe the state of a hydrogen atom in its ground and excited states.

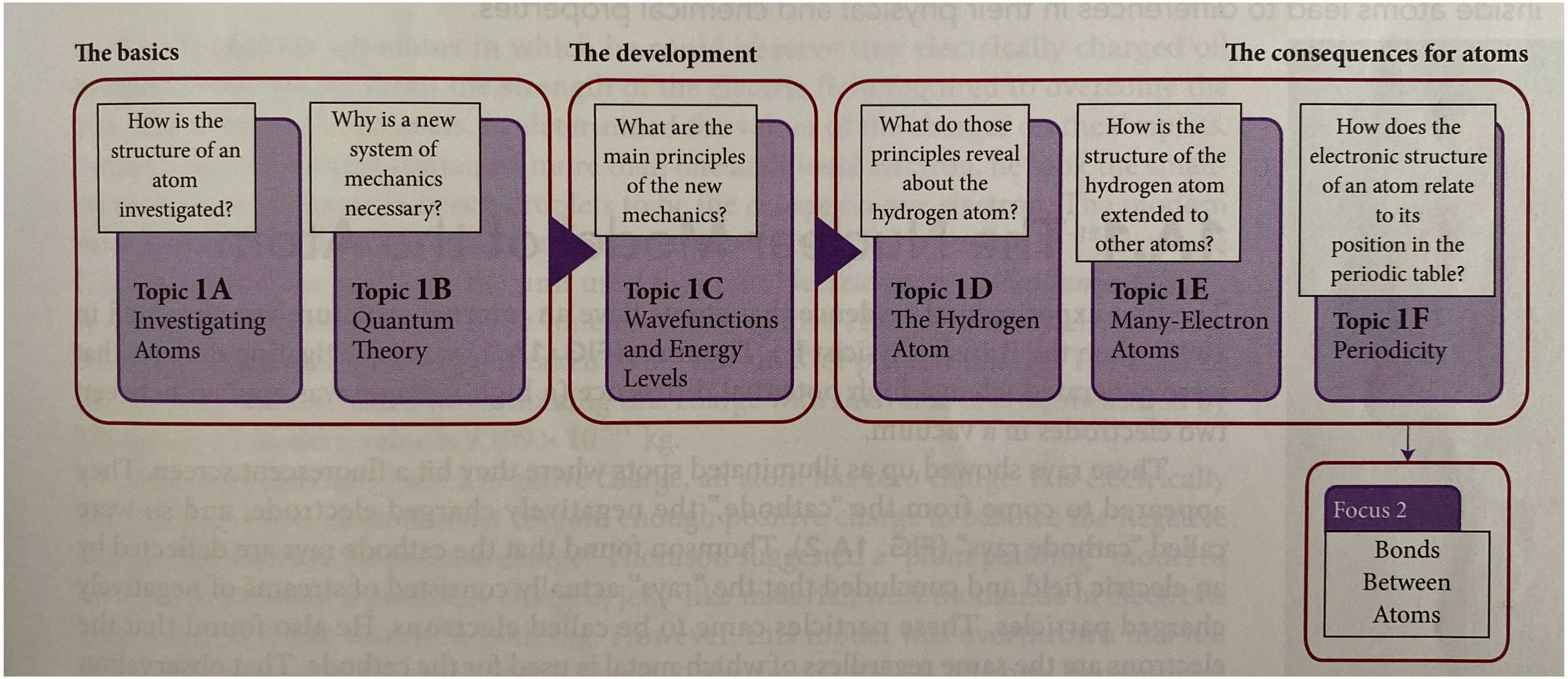
Summary: You have learned that an electron in a hydrogen atom is described by wavefunctions called atomic orbitals and that each orbital is specified by three quantum numers: n, l, and m_l. You now know that the shape and energy of a given orbital is found by solving the Schrödinger equation for an electron attracted to a nucleus. You also now know that transitions between the allowed energy levels account for the observed patterns of spectroscopic lines. You have also encountered the property of “electron spin” and know that electron spin may have either of two orientations.

	Particle in a box	**Hydrogen atom**
Dimension of space	1D	3D
Walls	Physical walls	No physical walls, and electrons are confined by pull of the nucleus
Quantization	Energy quantized	
Potential energy	Potential energy inside the box is zero	Potential energy governed by Coulomb potential
Wave function shape	Sinusoidal functions (sine or cosine)	Wave functions (called **orbitals**) are more complex, often **spherical** or **lobed** in shape (spherical harmonics), with both radial and angular components.
Quantum numbers	One quantum number, n, which represents the energy level and is related to the number of nodes in the wave function.	Three quantum numbers: $\boldsymbol{n}$: principal quantum number (energy level), $\boldsymbol{l}$: angular momentum quantum number (shape of the orbital), $\boldsymbol{m_l}$: magnetic quantum number (orientation of the orbital).
Degeneracy	No degeneracy: each energy level corresponds to one unique state.	Degeneracy in energy levels: for a given principal quantum number n, multiple different orbitals (characterized by l and m_l) have the same energy.
Boundary conditions	The wave function must go to **zero at the walls** of the box.	The wave function must go to zero at infinity, far from the nucleus.
Physical interpretation	The particle is **free** inside the box but cannot escape due to infinite potential at the walls.	The electron is **bound** to the nucleus due to the attractive Coulomb force, which confines the electron.

Many-Electron Atoms

Topic 1E

Overview Chapter 1 (Focus 1: Atoms)



The basics

How is the structure of an atom investigated?

Topic 1A Investigating Atoms

Why is a new system of mechanics necessary?

Topic 1B Quantum Theory

The development

What are the main principles of the new mechanics?

Topic 1C Wavefunctions and Energy Levels

The consequences for atoms

What do those principles reveal about the hydrogen atom?

Topic 1D The Hydrogen Atom

How is the structure of the hydrogen atom extended to other atoms?

Topic 1E Many-Electron Atoms

How does the electronic structure of an atom relate to its position in the periodic table?

Topic 1F Periodicity

Focus 2

Bonds Between Atoms

Topic 1E.1 Orbital Energies
Topic 1E.2 The Building-Up Principle

WHY DO YOU NEED TO KNOW THIS MATERIAL?

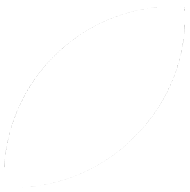
- The electronic structures of many-electron atoms account for the **form** of chemistry's all-important **periodic table**.

WHAT DO YOU NEED TO KNOW ALREADY?

- Description of atomic orbitals of hydrogen (**Topic 1D**), especially their radial dependence and their angular shapes.
- Electron has a property called spin.
- General structure of the periodic table (**Fundamentals B**)

Orbital energies

Topic 1E.1



1E.1 Orbital energies

Most atoms have more than one electron!

- A neutral atom other than hydrogen has more than one electron and is known as a **many-electron atom** (or polyelectron atom).
- In this Topic 1E, we will learn how the presence of more than one electron affects the **energies of atomic orbitals** and **how they are occupied**.

1E.1 Orbital energies

Most atoms have more than one electron

- As with hydrogen, electrons in many-electron atoms **occupy orbitals**. Two main differences:

1. **Nucleus is more highly charged** → attracts electrons more strongly → lowers otbital energies
2. **Electrons repel one another** → repulsion opposes nuclear attraction → raises orbital energies

1E.1 Orbital energies

Energies of hydrogen vs. many-electron atoms

Hydrogen atom	**Many-electron atom**
One electron	Many electrons
No electron-electron repulsion	Electron-electron repulsion
All orbitals within same shell are degenerate (have the same energy)	Electron-electron repulsion raises 2p-orbital energy (2s and 2p are not degenerate)
2s- and 2p-orbitals all have same energy	2p-orbital energy is higher than 2– orbital energy

1E.1 Orbital energies

Shielding

1E.1 Orbital energies

Shielding in different orbitals

- Extent of shielding depends on orbital shape
- **s-electrons**: high probability near nucleus → weakly shielded → feel stronger nuclear charge → lower energy.
- **p-electrons**: no probability at nucleus, less close approach → more shielding → feel weaker nuclear charge → higher energy.
- **d-electrons**: even less close approach → most shielding → weakest nuclear attraction → highest energy.

1E.1 Orbital energies

Radial distribution functions for s-, p-, and d-orbitals

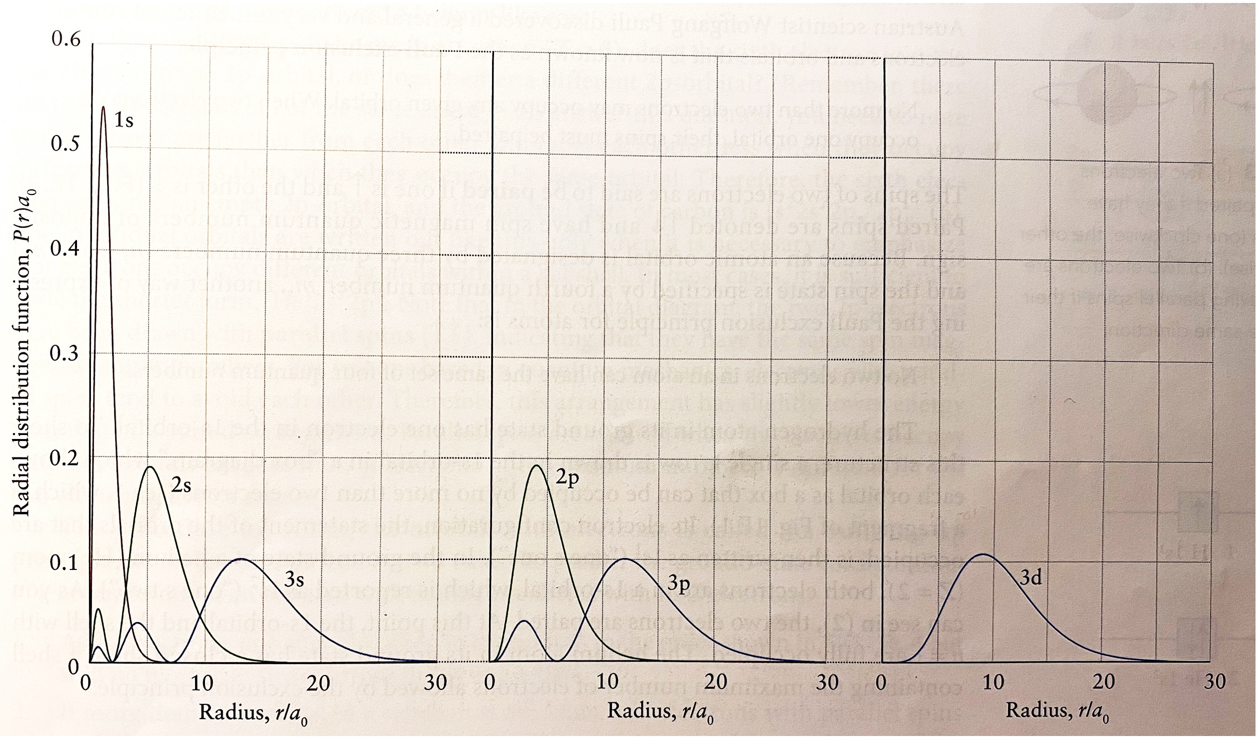


Figure 1E.2

- In the same shell: probability maxima occur at similar distances
- s > p > d in probability near nucleus

1E.1 Orbital energies

Effects of shielding can be large

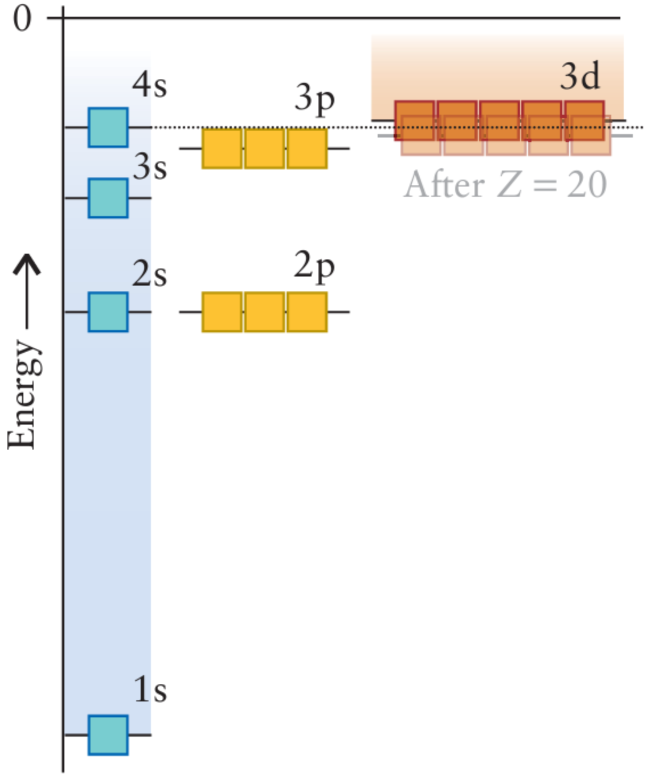


Figure 1E.1

- A 4s-electron has a much lower energy than a 3d-electron of the same atom.
- The **precise ordering** of the orbitals depends on the **number of electrons** in the atom as well as the effective nuclear charge experienced by the electrons (see next section)

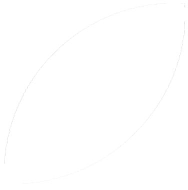
1E.1 Orbital energies

Summary

In a many-electron atom, because of the effects of penetration and shielding, the order of orbital energies in a given shell is s < p < d < f.

The Building-Up Principle

Topic 1E.2



1E.2 The building-up principle

Electronic structure of many-electron atoms

- Electronic structure of an atom determines chemical properties.
- **Electron configuration**: a list of all its occupied orbitals with the numbers of electrons that occupy each one.

1E.2 The building-up principle

The Pauli exclusion principle

- In 1925, Austrian scientist Wolfgang Pauli discovered a general and fundamental rule about electrons and orbitals:

No more than two electrons may occupy any given orbital. When two electrons do occupy one orbital, their spins must be paired.

- The two spins of electrons are paired if one is ↑ and ↓.
- Paired spins are denoted ↑↓ and have spin magnetic quantum numbers of opposite signs.

1E.2 The building-up principle

The Pauli exclusion principle

(a) The two spins of electrons are **paired** if one is ↑ and ↓. They have opposite signs (one clockwise, the other counterclockwise).

(b) Two electrons are classified as having **parallel** spins if their spins are in the same direction.

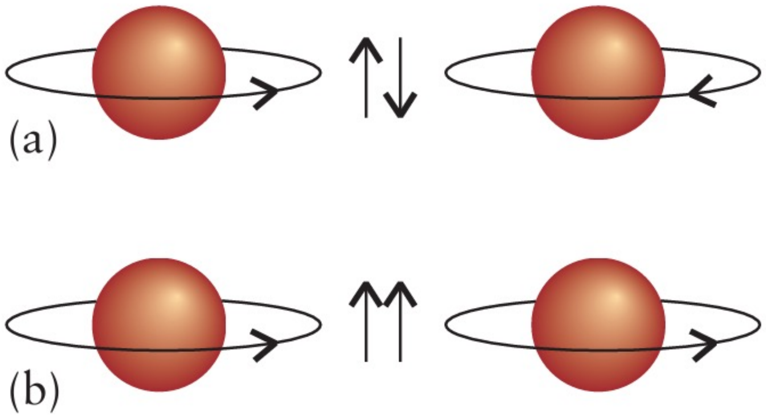
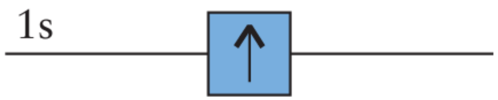


Figure 1E.3

1E.2 The building-up principle

Hydrogen and helium

- Hydrogen in the ground state: one electron in the 1s-orbital: **$1s^1$** (electron configuration)
- Box indicates space for two electrons.
- Helium in the ground state: **$1s^2$**
- Helium has a **fully occupied** 1s orbital
- Helium has **a closed shell:** a shell containing the maximum number of electrons allowed by the exclusion principle.



1 H $1s^1$

2 He $1s^2$

1E.2 The building-up principle

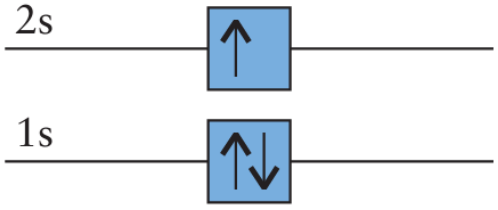
Core and valence electrons

- Filled, inner orbitals: **core**
- Outermost shell: **valence electrons**
- Core electrons are in lower-energy orbitals and tightly bound.
- The outermost electrons are used in the formation of chemical bonds (Topic 2A), and one theory of bond formation is called ***valence-bond*** **theory**, hence the name *valence electrons*.

1E.2 The building-up principle

Lithium (Li)

- Lithium (Z = 3) has three electrons: Two in 1s-, one in the 2s-orbital.
- Ground state of lithium: **$1s^2 2s^1$**
- Core for lithium: $1s^2$ = [He]
- With valence electrons: **$[He]2s^1$**
- Lithium only loses one electron when it forms compounds: Li^+ rather than Li^{2+} or Li^{3+}

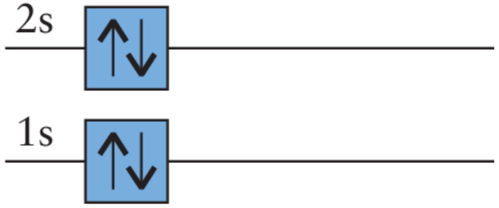


3 Li $1s^2 2s^1$, $[He]2s^1$

1E.2 The building-up principle

Beryllium (Be)

- Beryllium (Z = 4) has four electrons: Two in 1s-, two in the 2s-orbital.
- Ground state: **$1s^2 2s^2$**
- With valence electrons: **[He]$2s^2$**
- Be atom loses only valence shell electrons in chemical reactions: Be^{2+} ion

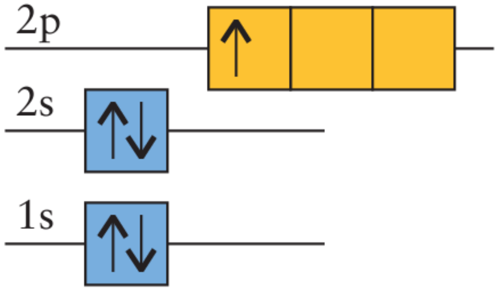


4 Be $1s^2 2s^2$, [He]$2s^2$

1E.2 The building-up principle

Boron (B)

- Boron (Z = 5) has five electrons. Two in 1s-, two in the 2s-orbital, one in the 2p-orbital.
- Ground state: $1s^2 2s^2 2p^1$
- With valence electrons: $[He]2s^2 2p^1$

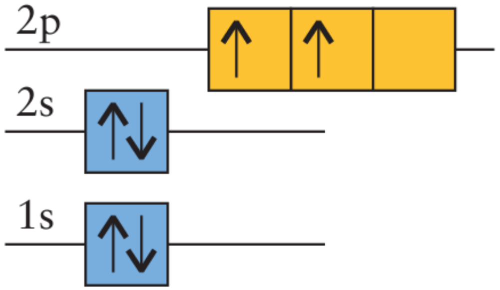


5 B $1s^2 2s^2 2p^1$, $[He]2s^2 2p^1$

1E.2 The building-up principle

Carbon (C)

- Carbon (Z = 6) has six electrons. Two in 1s-, two in the 2s-orbital, two in the 2p-orbital.
- Ground state: **$1s^2 2s^2 2p^2$**
- With valence electrons: **$[He]2s^2 2p^2$**
- **Decision**: paired or parallel electrons in p-orbitals?
- Electrons are farther from each other and repell each other less when they occupy different p-orbitals than when they occupy the same orbital:
- $1s^2 2s^2 2p_x^1 2p_y^1$
- Parallel spins!



6 C $1s^2 2s^2 2p^2$, $[He]2s^2 2p^2$

1E.2 The building-up principle

The building-up principle and Hund's rule

Two rules:

1. **Add *Z* electrons, one after the other, to the orbitals in the order shown in Fig. 1E.4 (next slide) but with no more than two electrons in any one orbital.**
2. **If more than one orbital in a subshell is available, add electrons with parallel spins to different orbitals of that subshell rather than pairing two electrons in one of the orbitals. (Hund's rule, after German spectroscopist Friedrich Hund)**

- **Configuration** of the atom at the **lowest total energy**: maximizes the attraction of the electrons to nucleus and minimizes their repulsion by one another.

1E.2 The building-up principle

The periodic table

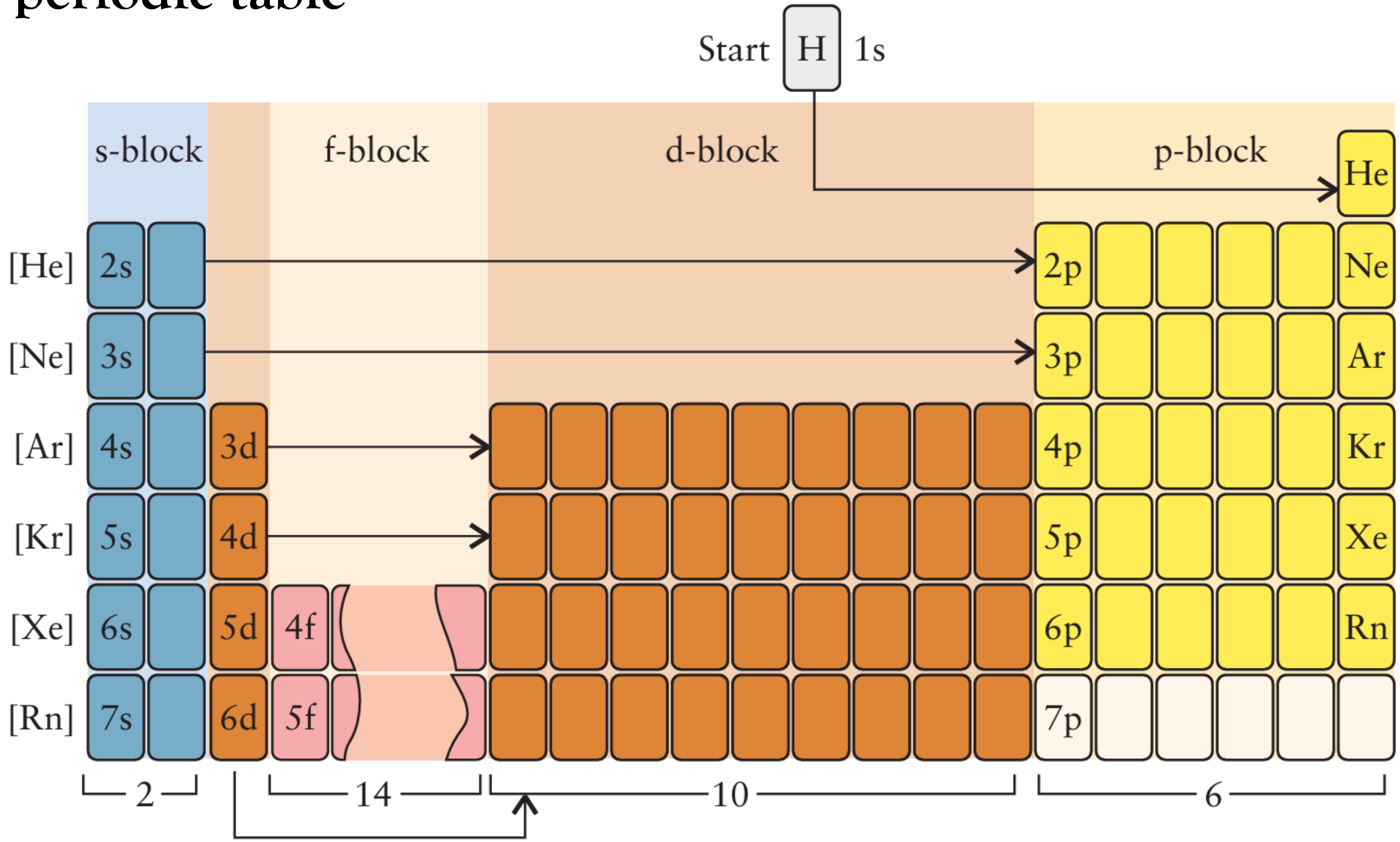


Figure 1E.4

1E.2 The building-up principle

Excited state electron configurations

- An atom with electrons in higher energy states than predicted by the building-up principle is said to be in **an excited state**.
- For example: [He]$2s^1 2p^3$ represents an excited state of a carbon atom.
- An excited state is **unstable** and quickly **relaxes** back to the ground state as the electron returns to an orbital that restores the atom to a lower energy and **emits a photon.**

1E.2 The building-up principle

The underlying organization: Periods

- Rows are called **periods**
- All the atoms of the **main-group elements** in a given period have a valence shell with the same principal quantum number (equal to period number)
- E.g. valence shell of elements in period 2 (lithium to neon) is the shell with n = 2.
- All atoms in a given period have the **same type of core** but **different numbers of valence electrons**.
- Period 2: elements that have a helium-like $1s^2$ core, denoted [He]
- Period 3: elements with a neon-like $1s^2\,2s^22p^6$ core, denoted [Ne]

1E.2 The building-up principle

The underlying organization: Groups

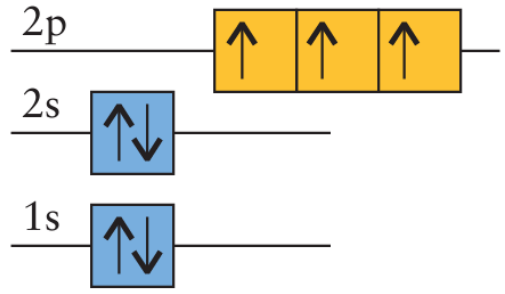
- **Columns** are **groups**: main groups 1, 2, 13-18
- Same group, analogous valence electron configurations, differ only in value of *n*
- E.g. all members of Group 1 have the valence configuration ns^1
- All members of Group 14 have the valence configuration ns^2np^2
- **Similar electron configurations** give the elements in a group **similar chemical properties**

1E.2 The building-up principle

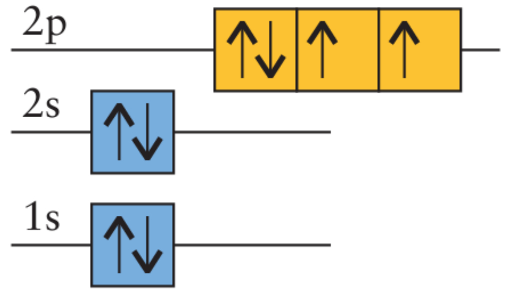
Nitrogen and oxygen

- **Nitrogen** (Z = 7) has seven electrons. Two in 1s-, two in the 2s-orbital, three in the 2p-orbital.

- **Oxygen** (Z = 8) has eight electrons. Two in 1s-, two in the 2s-orbital, four in the 2p-orbital.



7 N $1s^2 2s^2 2p^3$, $[He]2s^2 2p^3$

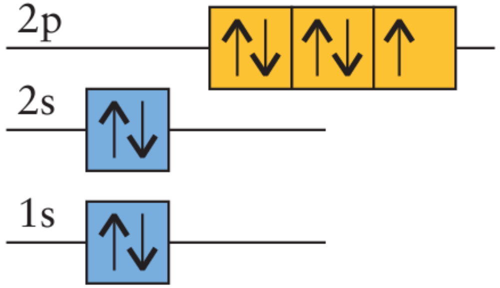


8 O $1s^2 2s^2 2p^4$, $[He]2s^2 2p^4$

1E.2 The building-up principle

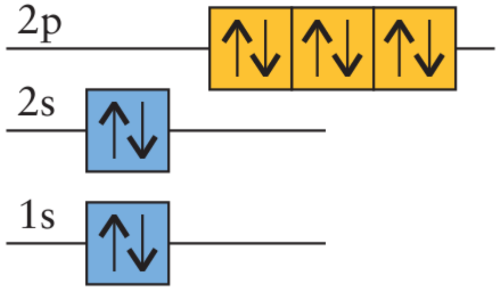
Fluorine and neon

- **Fluorine** (Z = 9) has nine electrons. Two in 1s-, two in the 2s-orbital, five in the 2p-orbital.
- Only one unpaired electron



9 F $1s^2 2s^2 2p^5$, [He]$2s^2 2p^5$

- **Neon** (Z = 10) has ten electrons. Two in 1s-, two in the 2s-orbital, six in the 2p-orbital.
- No unpaired electrons (complete shell *n* = 2)



10 Ne $1s^2 2s^2 2p^6$, [He]$2s^2 2p^6$

1E.2 The building-up principle

After the 3p orbitals are filled…

- s- and p-orbitals of the shell with n = 3 full at **argon**: [Ne]$3s^2 3p^6$
- **Next orbital 4s** (not 3d!) because s-electrons penetrate through the inner-core electrons to a greater extent than p- or d-electrons → lowers energy of 4s orbital
- Potassium [Ar]$4s^1$ and calcium [Ar]$4s^2$ are next
- [Ar] denotes $1s^2 2s^2 2p^6 3s^2 3p^6$ core
- Then 3d-orbitals are filled

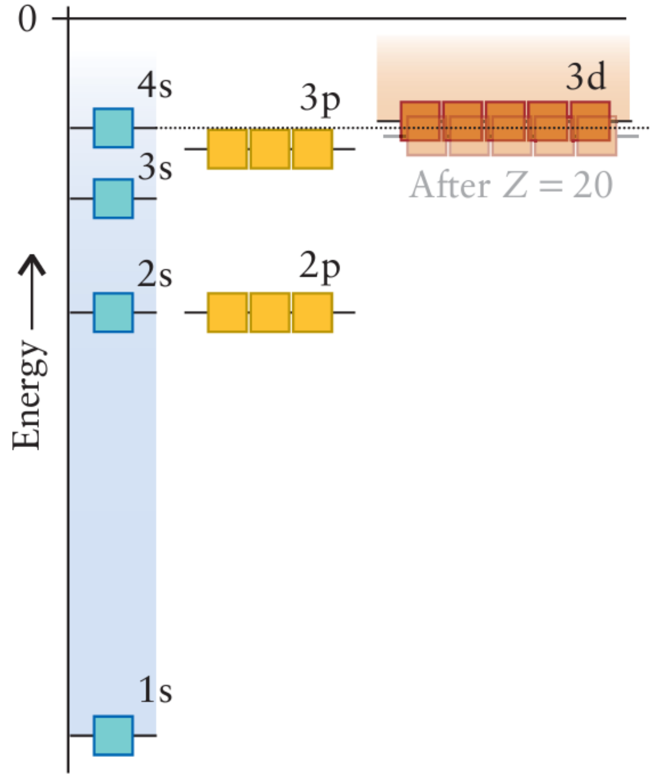


Figure 1E.1

1E.2 The building-up principle

After the 4s orbital is filled…

- Change of rhythm: **3d orbitals**
- From $Z = 21$ to $Z = 30$ (scandium to zinc)
- Scandium ($Z = 21$): [Ar]$3d^1 4s^2$
- Titanium ($Z = 22$): [Ar]$3d^2 4s^2$
- Note: 3d-orbitals are written before 4s-orbitals because they become lower in energy compared to 4s **when filled with electrons**

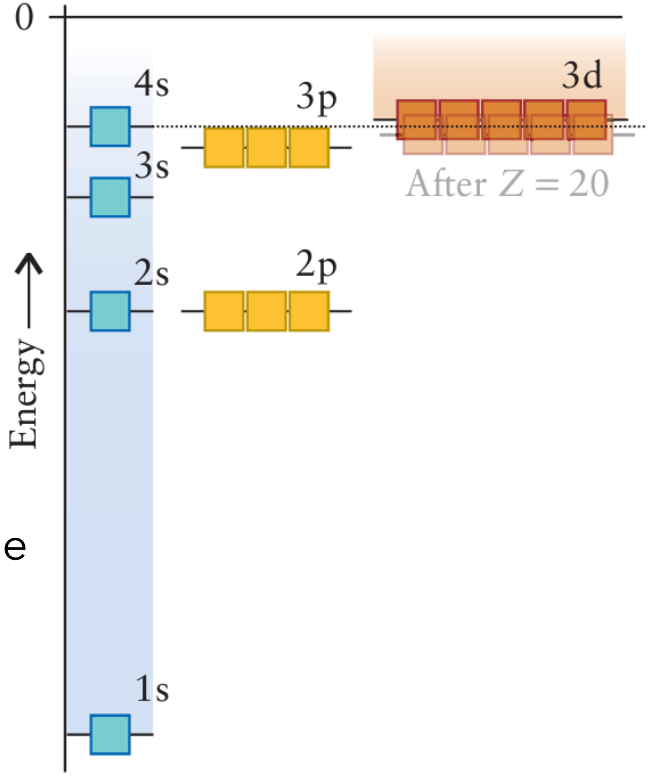


Figure 1E.1

1E.2 The building-up principle

After the 4s orbital is filled…

- Successive electrons are added to d-electrons as *Z* increases.

With two exceptions:

- Chromium (Z = 24): [Ar]$3d^5 4s^1$ instead of [Ar]$3d^4 4s^2$
- Copper (Z = 29): [Ar]$3d^{10} 4s^1$ instead of [Ar]$3d^9 4s^2$
- **Half-complete subshell configuration** d^5 and **complete subshell configuration** d^{10} have a lower energy

1E.2 The building-up principle

After the 3d orbitals are filled…

- Note: Electron configurations are written **in order of increasing energy, not in order of filling**. For example, scandium is written [Ar]$3d^1 4s^2$ and not [Ar]$4s^2 3d^1$
- 4p-orbitals are next (see periodic table!)
- Germanium: [Ar]$3d^{10} 4s^2 4p^2$
- Arsenic: [Ar]$3d^{10} 4s^2 4p^3$
- Fourth period contains **18 elements**: 4s- and 4p-orbitals with 8 electrons and 3d-orbitals with 10 electrons
- Period four is the first **long period** of the periodic table

TOOLBOX 1.1 HOW TO PREDICT THE GROUND-STATE ELECTRON CONFIGURATION OF AN ATOM

CONCEPTUAL BASIS

Electrons occupy orbitals in such a way as to minimize the total energy of an atom by maximizing attractions and minimizing repulsions in accord with the Pauli exclusion principle and Hund's rule.

PROCEDURE

We use the following rules of the building-up principle to assign a ground-state configuration to a neutral atom of an element with atomic number Z:

1. Add Z electrons, one after the other, to the orbitals in the order shown in Figs. 1.41 and 1.44 but with no more than two electrons in any one orbital (the Pauli exclusion principle).
2. If more than one orbital in a subshell is available, add electrons to different orbitals of the subshell before doubly occupying any of them (Hund's rule).
3. Write the labels of the orbitals in order of increasing energy, with a superscript that gives the number of electrons in that orbital. The configuration of a filled shell is represented by the symbol of the noble gas having that configuration, as in [He] for $1s^2$.
4. When drawing a box diagram, show the electrons in different orbitals of the same subshell with parallel spins; electrons sharing an orbital have paired spins.

In most cases this procedure gives the ground-state electron configuration of an atom, the arrangement with the lowest energy. Any arrangement other than the ground state corresponds to an excited state of the atom. Note that we can use the structure of the periodic table to predict the electron configurations of most elements once we realize which orbitals are being filled in each block of the periodic table (see Fig. 1.44).

A useful shortcut for atoms of elements with large numbers of electrons is to write the valence electron configuration from the group number, which gives the number of valence electrons in the ground state of the atom, and the period number, which gives the value of the principal quantum number of the valence shell. The core consists of the preceding noble-gas configuration together with any completed d- and f-subshells.

Example 1.10 shows how these rules are applied.

1E.2 The building-up principle

After the 4p orbitals are filled…

- Period 5: **5s-orbital** is filled next, followed by **4d-orbitals**
- As in Period 4, the energies of the 4d-orbitals fall below that of the 5s-orbitals after two electrons have been accommodated in the 5s-orbital
- Similar effect is seen in Period 6:
- Cerium: $[Xe]4f^{1}5d^{1}6s^{2}$
- Electrons then continue to occupy the seven 4f-orbitals, which are complete after 14 electrons have been added, at ytterbium, $[Xe]4f^{14}5d^{10}6s^{2}6p^{1}$

PERIODIC TABLE OF THE ELEMENTS

Period \ Group	1 I IA	2 II IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 III IIIA	14 IV IVA	15 V VA	16 VI VIA	17 VII VIIA	18 VIII VIIA
1	**1 H** hydrogen 1.0079 $1s^1$																	**2 He** helium 4.00 $1s^2$
2	**3 Li** lithium 6.94 $2s^1$	**4 Be** beryllium 9.01 $2s^2$											**5 B** boron 10.81 $2s^22p^1$	**6 C** carbon 12.01 $2s^22p^2$	**7 N** nitrogen 14.01 $2s^22p^3$	**8 O** oxygen 16.00 $2s^22p^4$	**9 F** fluorine 19.00 $2s^22p^5$	**10 Ne** neon 20.18 $2s^22p^6$
3	**11 Na** sodium 22.99 $3s^1$	**12 Mg** magnesium 24.31 $3s^2$											**13 Al** aluminum 26.98 $3s^23p^1$	**14 Si** silicon 28.09 $3s^23p^2$	**15 P** phosphorus 30.97 $3s^23p^3$	**16 S** sulfur 32.06 $3s^23p^4$	**17 Cl** chlorine 35.45 $3s^23p^5$	**18 Ar** argon 39.95 $3s^23p^6$
4	**19 K** potassium 39.10 $4s^1$	**20 Ca** calcium 40.08 $4s^2$	**21 Sc** scandium 44.96 $3d^14s^2$	**22 Ti** titanium 47.87 $3d^24s^2$	**23 V** vanadium 50.94 $3d^34s^2$	**24 Cr** chromium 52.00 $3d^54s^1$	**25 Mn** manganese 54.94 $3d^54s^2$	**26 Fe** iron 55.84 $3d^64s^2$	**27 Co** cobalt 58.93 $3d^74s^2$	**28 Ni** nickel 58.69 $3d^84s^2$	**29 Cu** copper 63.55 $3d^{10}4s^1$	**30 Zn** zinc 65.41 $3d^{10}4s^2$	**31 Ga** gallium 69.72 $4s^24p^1$	**32 Ge** germanium 72.64 $4s^24p^2$	**33 As** arsenic 74.92 $4s^24p^3$	**34 Se** selenium 78.96 $4s^24p^4$	**35 Br** bromine 79.90 $4s^24p^5$	**36 Kr** krypton 83.80 $4s^24p^6$
5	**37 Rb** rubidium 85.47 $5s^1$	**38 Sr** strontium 87.62 $5s^2$	**39 Y** yttrium 88.91 $4d^15s^2$	**40 Zr** zirconium 91.22 $4d^25s^2$	**41 Nb** niobium 92.91 $4d^45s^1$	**42 Mo** molybdenum 95.94 $4d^55s^1$	**43 Tc** technetium (98) $4d^55s^2$	**44 Ru** ruthenium 101.07 $4d^75s^1$	**45 Rh** rhodium 102.90 $4d^85s^1$	**46 Pd** palladium 106.42 $4d^{10}$	**47 Ag** silver 107.87 $4d^{10}5s^1$	**48 Cd** cadmium 112.41 $4d^{10}5s^2$	**49 In** indium 114.82 $5s^25p^1$	**50 Sn** tin 118.71 $5s^25p^2$	**51 Sb** antimony 121.76 $5s^25p^3$	**52 Te** tellurium 127.60 $5s^25p^4$	**53 I** iodine 126.90 $5s^25p^5$	**54 Xe** xenon 131.29 $5s^25p^6$
6	**55 Cs** cesium 132.91 $6s^1$	**56 Ba** barium 137.33 $6s^2$	**57 La** lanthanum 138.91 $5d^16s^2$	**72 Hf** hafnium 178.49 $5d^26s^2$	**73 Ta** tantalum 180.95 $5d^36s^2$	**74 W** tungsten 183.84 $5d^46s^2$	**75 Re** rhenium 186.21 $5d^56s^2$	**76 Os** osmium 190.23 $5d^66s^2$	**77 Ir** iridium 192.22 $5d^76s^2$	**78 Pt** platinum 195.08 $5d^96s^1$	**79 Au** gold 196.97 $5d^{10}6s^1$	**80 Hg** mercury 200.59 $5d^{10}6s^2$	**81 Tl** thallium 204.38 $6s^26p^1$	**82 Pb** lead 207.2 $6s^26p^2$	**83 Bi** bismuth 208.98 $6s^26p^3$	**84 Po** polonium (209) $6s^26p^4$	**85 At** astatine (210) $6s^26p^5$	**86 Rn** radon (222) $6s^26p^6$
7	**87 Fr** francium (223) $7s^1$	**88 Ra** radium (226) $7s^2$	**89 Ac** actinium (227) $6d^17s^2$	**104 Rf** rutherfordium (261) $6d^27s^2$	**105 Db** dubnium (262) $6d^37s^2$	**106 Sg** seaborgium (266) $6d^47s^2$	**107 Bh** bohrium (264) $6d^57s^2$	**108 Hs** hassium (267) $6d^67s^2$	**109 Mt** meitnerium (268) $6d^77s^2$	**110 Ds** darmstadtium (271) $6d^87s^2$	**111 Rg** roentgenium (272) $6d^{10}7s^1$	112*	113	114	115	116	117	118

Lanthanoids (lanthanides) 6	**58 Ce** cerium 140.12 $4f^15d^16s^2$	**59 Pr** praseodymium 140.91 $4f^36s^2$	**60 Nd** neodymium 144.24 $4f^46s^2$	**61 Pm** promethium (145) $4f^56s^2$	**62 Sm** samarium 150.36 $4f^66s^2$	**63 Eu** europium 151.96 $4f^76s^2$	**64 Gd** gadolinium 157.25 $4f^75d^16s^2$	**65 Tb** terbium 158.93 $4f^96s^2$	**66 Dy** dysprosium 162.50 $4f^{10}6s^2$	**67 Ho** holmium 164.93 $4f^{11}6s^2$	**68 Er** erbium 167.26 $4f^{12}6s^2$	**69 Tm** thulium 168.93 $4f^{13}6s^2$	**70 Yb** ytterbium 173.04 $4f^{14}6s^2$	**71 Lu** lutetium 174.97 $5d^16s^2$
Actinoids (actinides) 7	**90 Th** thorium 232.04 $6d^27s^2$	**91 Pa** protactinium 231.04 $5f^26d^17s^2$	**92 U** uranium 238.03 $5f^36d^17s^2$	**93 Np** neptunium (237) $5f^46d^17s^2$	**94 Pu** plutonium (244) $5f^67s^2$	**95 Am** americium (243) $5f^77s^2$	**96 Cm** curium (247) $5f^76d^17s^2$	**97 Bk** berkelium (247) $5f^97s^2$	**98 Cf** californium (251) $5f^{10}7s^2$	**99 Es** einsteinium (252) $5f^{11}7s^2$	**100 Fm** fermium (257) $5f^{12}7s^2$	**101 Md** mendelevium (258) $5f^{13}7s^2$	**102 No** nobelium (259) $5f^{14}7s^2$	**103 Lr** lawrencium (262) $6d^17s^2$

Molar masses (atomic weights) quoted to the number of significant figures given here can be regarded as typical of most naturally occurring samples.

*The names of the elements 112 and higher have not yet been determined; both 112 and 114 have been confirmed.

1E.2 The building-up principle

Example 1E.1 Predicting the ground-state electron configuration of a heavy atom

- (a) vanadium
- (b) lead

1E.2 The building-up principle

Summary

The ground-state electron configuration of an atom is predicted by using the building-up principle in conjuction with Fig. 1E.1, the Pauli exclusion principle, and Hund's rule.

The skills you have mastered are the ability to

- ☐ Describe the factors affecting the energy of an electron in a many-electron atom.
- ☐ Write the ground-state electron configuration for an element.

Summary: You have learned that the structures of many-electron atoms are explained by the systematic occupation of orbitals by electrons, with the order determined by the effects of penetration and shielding in conjunction with the Pauli exclusion principle. The building-up principle is reflected in and in a sense accounts for the general structure of the periodic table.