7.8 Electron Spin

- The electron can be viewed as a ball of spinning charge – has a magnetic moment
- The magnetic moment is quantized – only two orientations of the spin are allowed in a magnetic field

⇒**Spin magnetic quantum number (m_s) –** two possible values of m_s **(+1/2 and -1/2)**

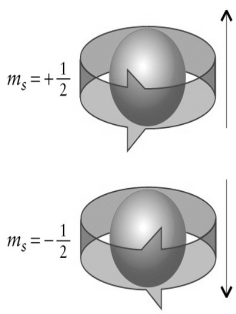

Fig. 7.25

7.9 The Electronic Structure of Hydrogen

- Electronic energy depends only on the principal quantum number (***n***) – all orbitals in a given shell have the same energy
- **Ground state** of the **H** atom – the lowest energy state (the electron occupies the **1s** orbital: ***n*=1, *l*=0, m_l=0, m_s=+1/2 or -1/2**)
- **Excited states** – higher energy states (***n*>1**)
- **Ionization energy (*I*)** – energy required to remove the electron from the atom in its ground state

$$I = E_\infty - E_1 = 0 - (-hR_H) = hR_H$$

The Electronic Structures of Multielectron Atoms

- Only approximate solutions of Schrödinger’s equation are available
- Electron-electron interactions are important
- The same four quantum numbers (***n, l,*** m_l and m_s) are used as for hydrogen

7.10 Orbital Energies

- Orbital energies depend on both ***n*** and ***l***

$$n\uparrow \rightarrow E\uparrow \qquad l\uparrow \rightarrow E\uparrow$$

- Orbitals in different subshells of a given principal shell have different energies

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- Electrons are attracted by the nucleus and repelled by each other
- **Electron shielding –** the inner electrons shield the outer electrons from the nuclear charge
- **Effective nuclear charge** (Z_{eff}) – smaller than the actual nuclear charge (***Z***)

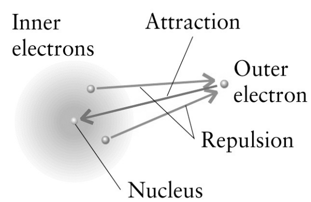


Fig. 6.27

- Electrons on orbitals in different subshells of a given shell are shielded to a different extent depending on their penetration to the nucleus
  - **s-orbitals** – high density at the nucleus, greatest penetration to the nucleus (least shielded, highest Z_{eff}, lowest energy)
  - **p-orbitals** – node at the nucleus, less penetration than the s-orbitals (more effectively shielded, larger Z_{eff}, higher energy)
  - penetration decreases with increasing ***l*** (more effective shielding, larger Z_{eff}, higher energy)
  - energy order of the subshells in a given shell:

s<p<d<f<g ...

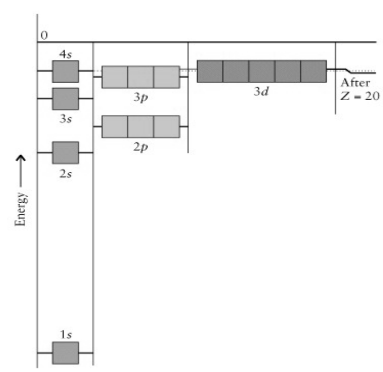
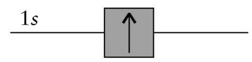


Fig. 6.28

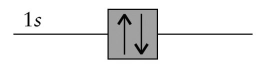
7.11 The Building-Up Principle

- The **Pauli exclusion principle** – no two electrons in an atom can have the same set of four quantum numbers
  - ⇒**each orbital can hold no more than two electrons and they must have opposite spins (paired spins, ↑↓)**
- **Building-up principle** – as new electrons are added to the atom, they are placed in the lowest energy available orbital (minimization of the total energy of the atom)

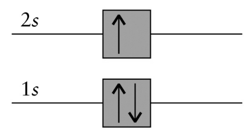
- **Electron configuration** – a list of the occupied orbitals and the number of electrons on them
- **Orbital diagrams**



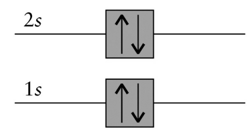
1 H $1s^1$



2 He $1s^2$

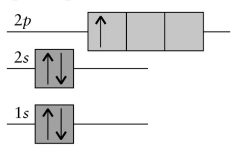


3 Li $1s^22s^1$

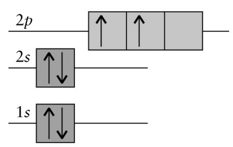


4 Be $1s^22s^2$

- **Closed shell** configuration – completely filled principal shell (**He** → $\mathbf{1s^2}$)



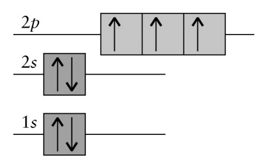
5 B $1s^22s^22p^1$



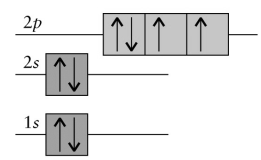
6 C $1s^22s^22p^2$

- **Degenerate orbitals** – have equal energies (all orbitals in a subshell)
- **Hund's rule** – in filling degenerate orbitals, electrons enter the empty orbitals having identical spins before pairing in one of them (minimization of the repulsion between the electrons)

- **Valence shell** – the outermost occupied shell
- **Valence electrons** – important in chemical reactions
- **Core electrons** – inner shells (abbreviated in the electron configurations, $\mathbf{1s^2}$ → abbreviated as **[He]**)



7 N $1s^22s^22p^3$



8 O $1s^22s^22p^4$, $[He]2s^22p^4$

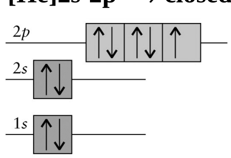
Example: Predict the electron configurations of **F** and **Ne**.

orbital order: 1s, 2s, 2p, 3s, 3p, …

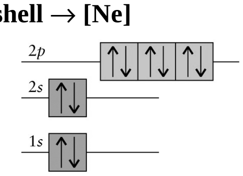
F (Z=9, 9 e^-) → $1s^22s^22p^5$ → $[He]2s^22p^5$

Ne (Z=10, 10 e^-) → $1s^22s^22p^6$ → $[He]2s^22p^6$

$[He]2s^22p^6$ → closed shell → [Ne]



9 F $1s^22s^22p^5$, $[He]2s^22p^5$



10 Ne $1s^22s^22p^6$, $[He]2s^22p^6$