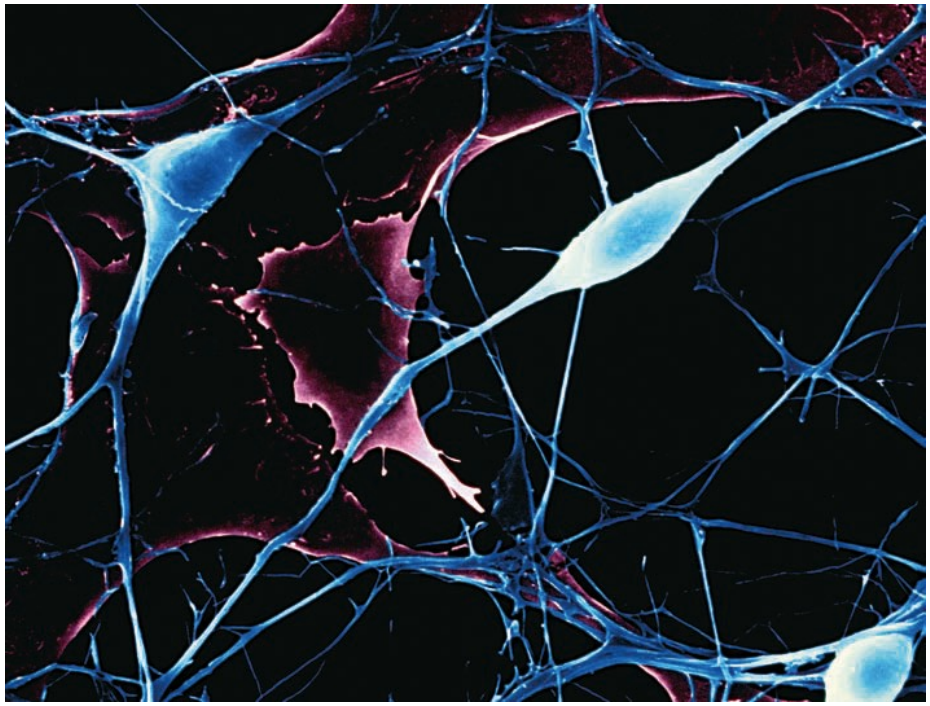
PK Chemie

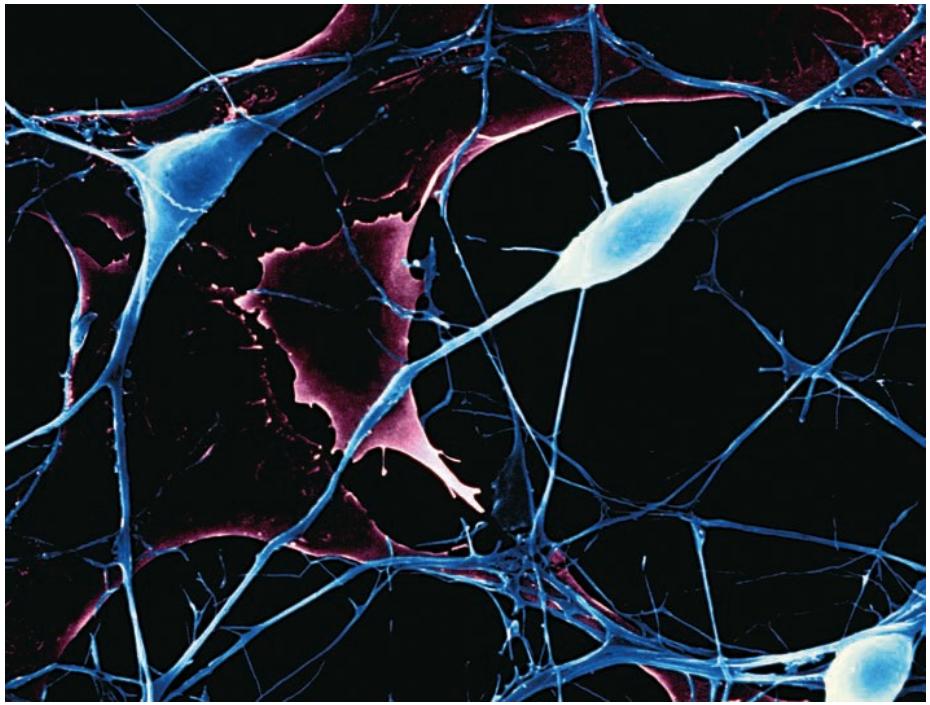
Elektronen in Atomen

Chemische Verbindungen



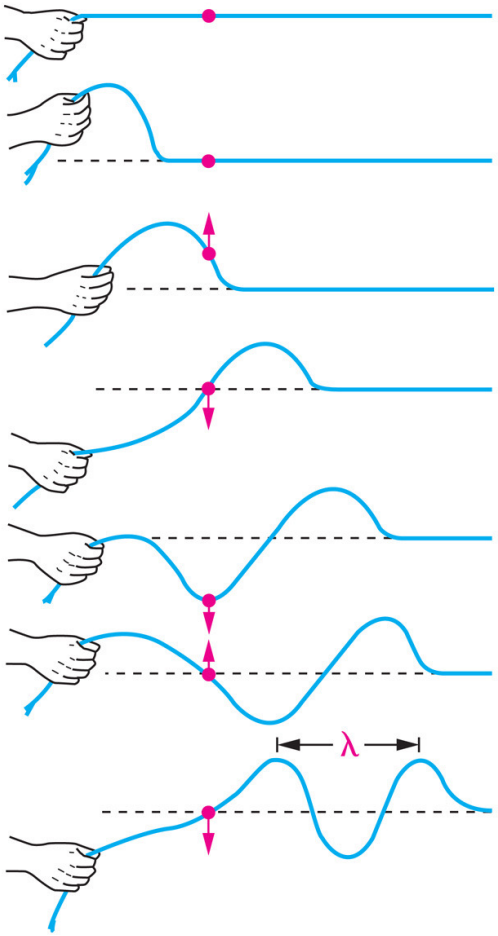
INHALTE

Chemical Compounds



	INHALTE
8	Elektron-Spin
9	Atome mit mehreren Elektronen
10	Elektronenkonfigurationen
11	Elektronenkonfigurationen und das PSE

Elektromagnetische Strahlung

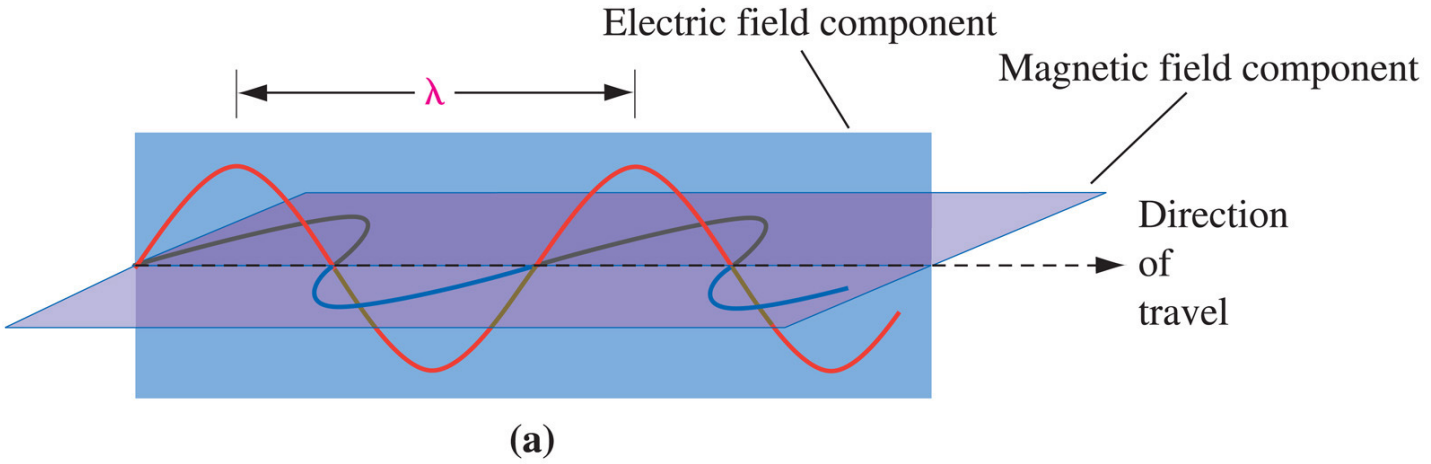


Elektrische und magnetische Felder bewegen sich wie Wellen durch den Raum.

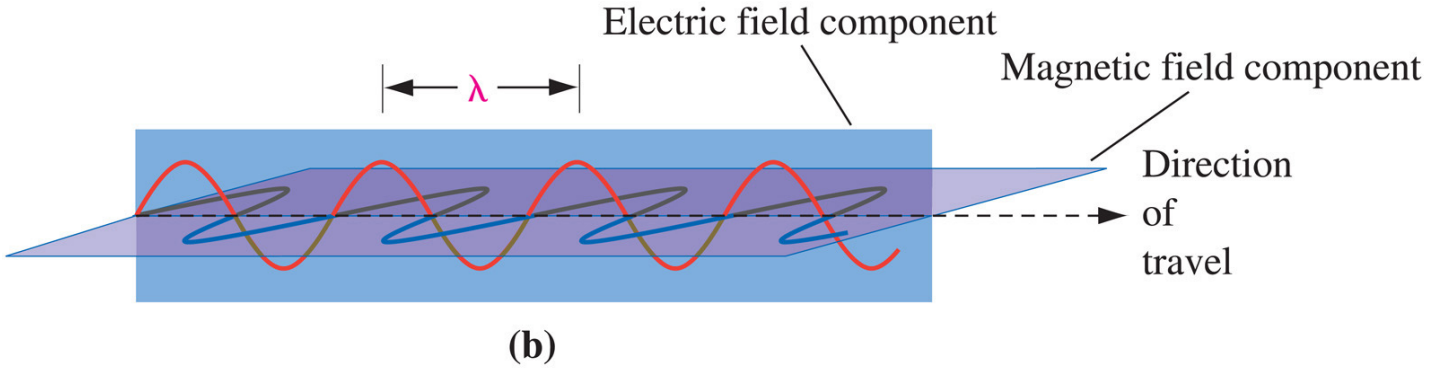
Eine Welle übermittelt Energie.

▲ **Wellenbewegung eines Seiles**

Kleines ν



Grosses ν



▲

Elektromagnetische Wellen

Frequenz, Wellenlänge und Geschwindigkeit der elektromagnetischen Strahlung

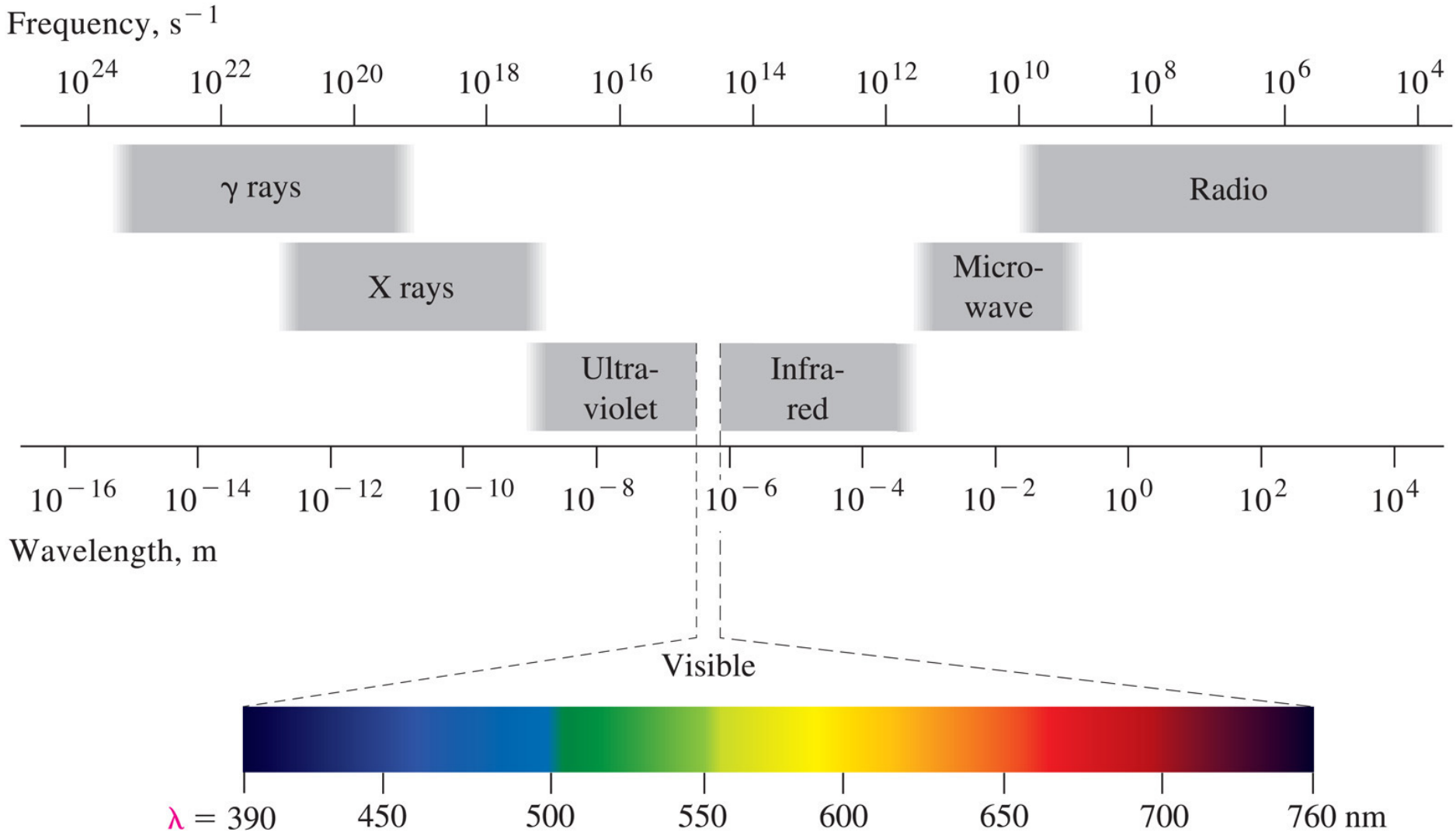
Frequenz (ν) in Hertz—Hz oder s^{-1}.

Wellenlänge (λ) in Meter—m.

cm	**μm**	**nm**	**Å**	**p**
(10^{-2} m)	(10^{-6} m)	(10^{-9} m)	(10^{-10} m)	(10^{-12} m)

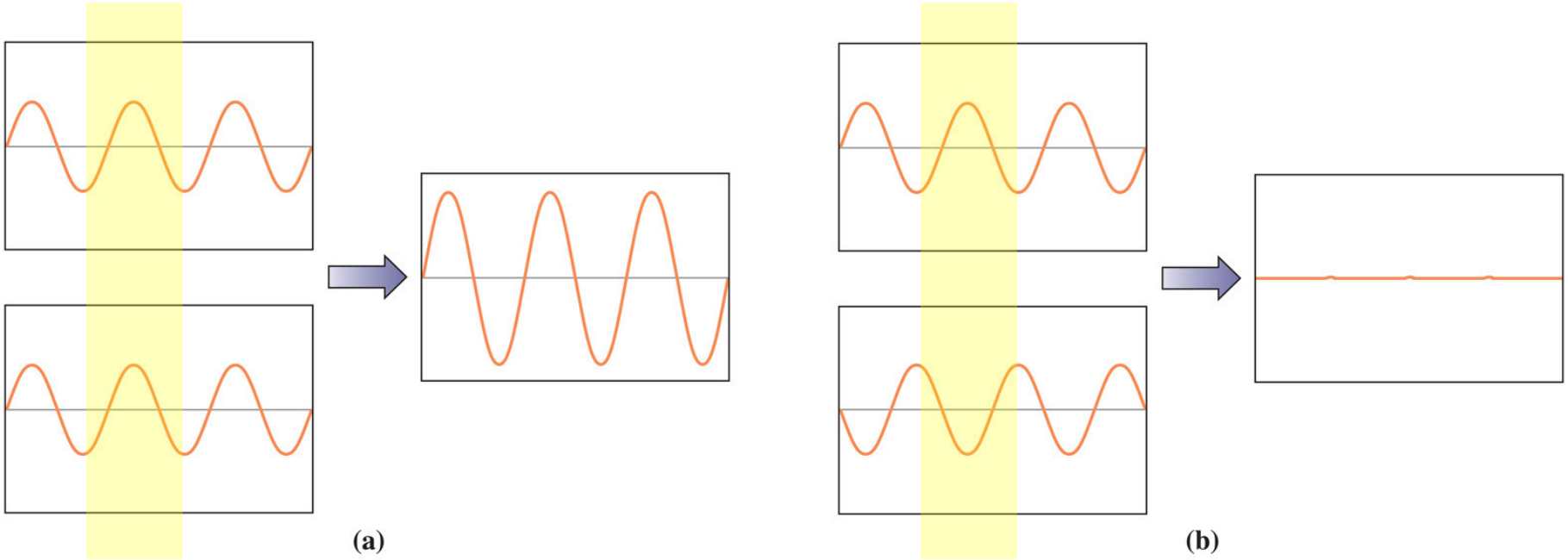
Licht-Geschwindigkeit: (c)—2.997925×10^8 m s^{-1}.

$$c = \lambda \nu \qquad \lambda = c/\nu \qquad \nu = c/\lambda$$



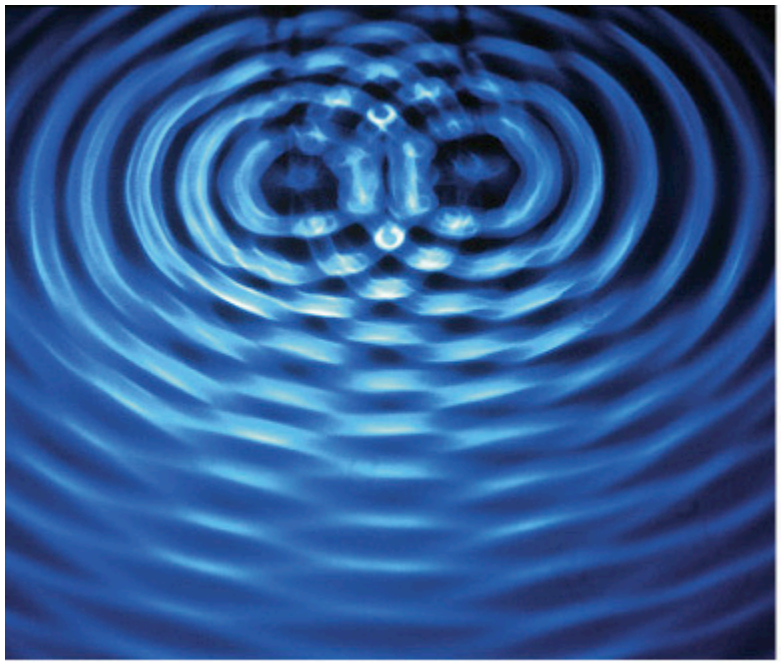
▲

Elektromagnetisches Spektrum

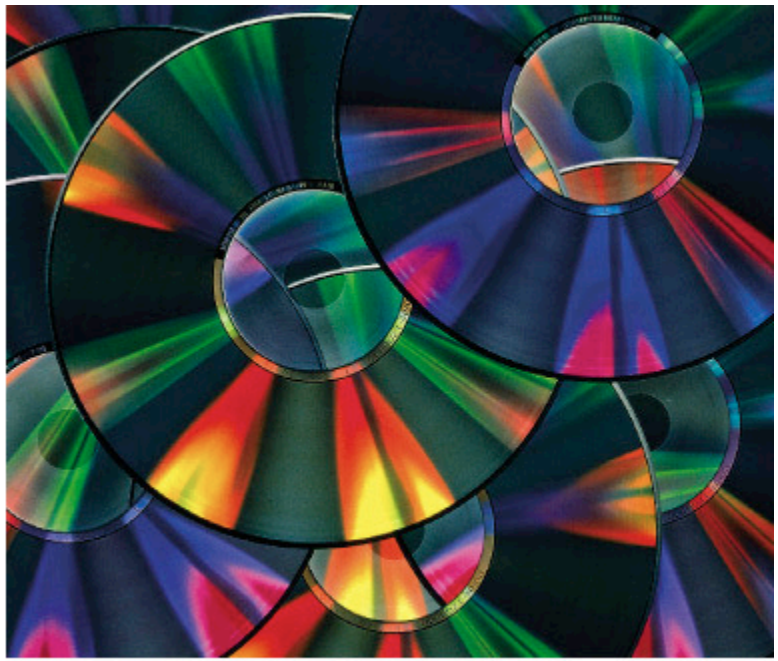


▲

Interferenz zweier überlappender Wellen



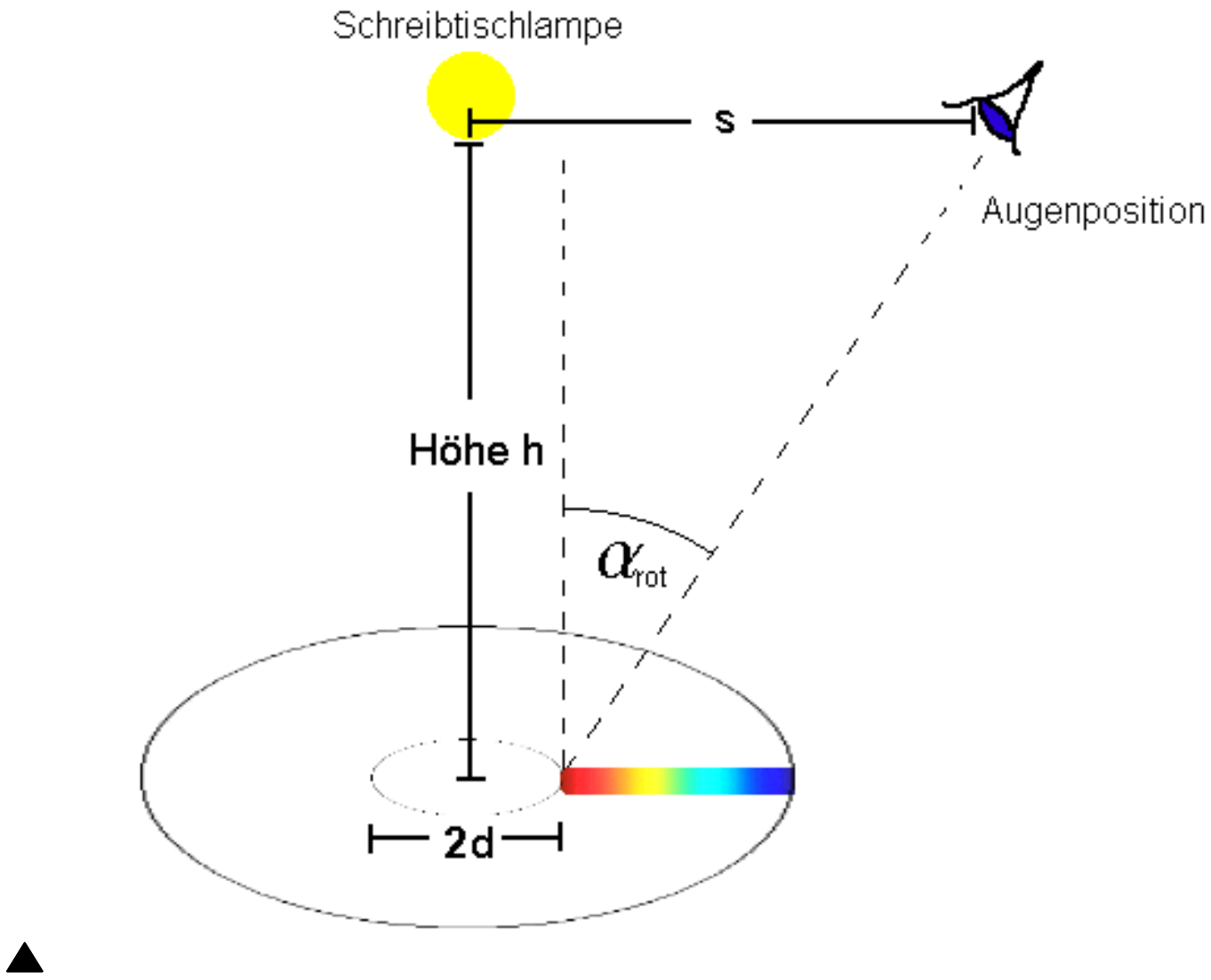
(a)

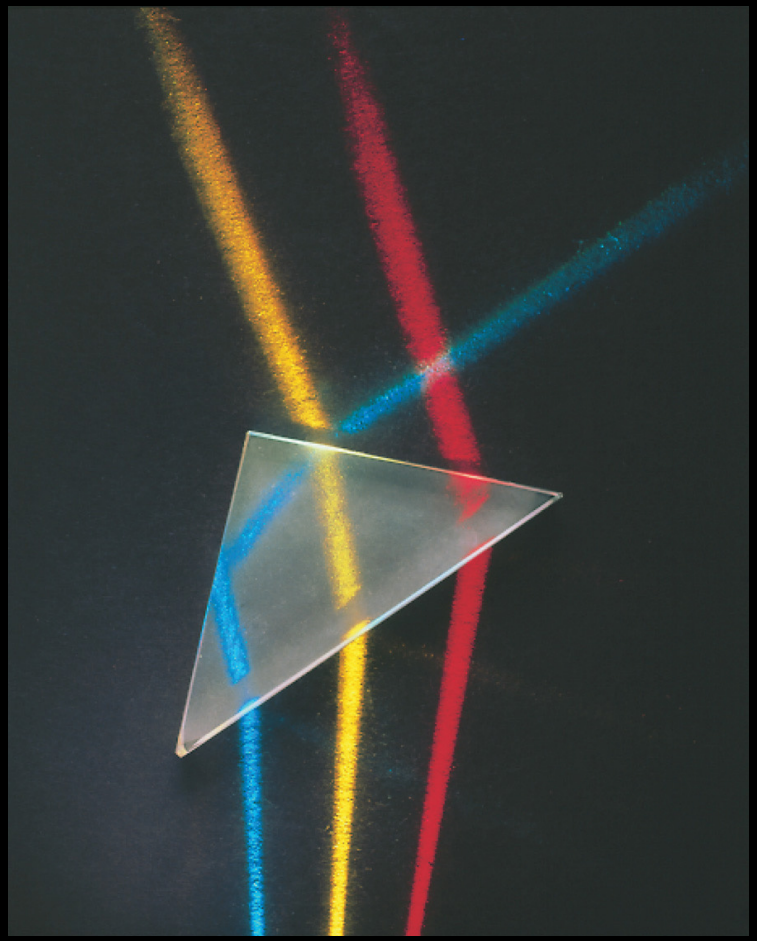


(b)

▲

Interferenz-Beispiele


Schreibtischlampe
s
Augenposition
Höhe h
α_{rot}
2d

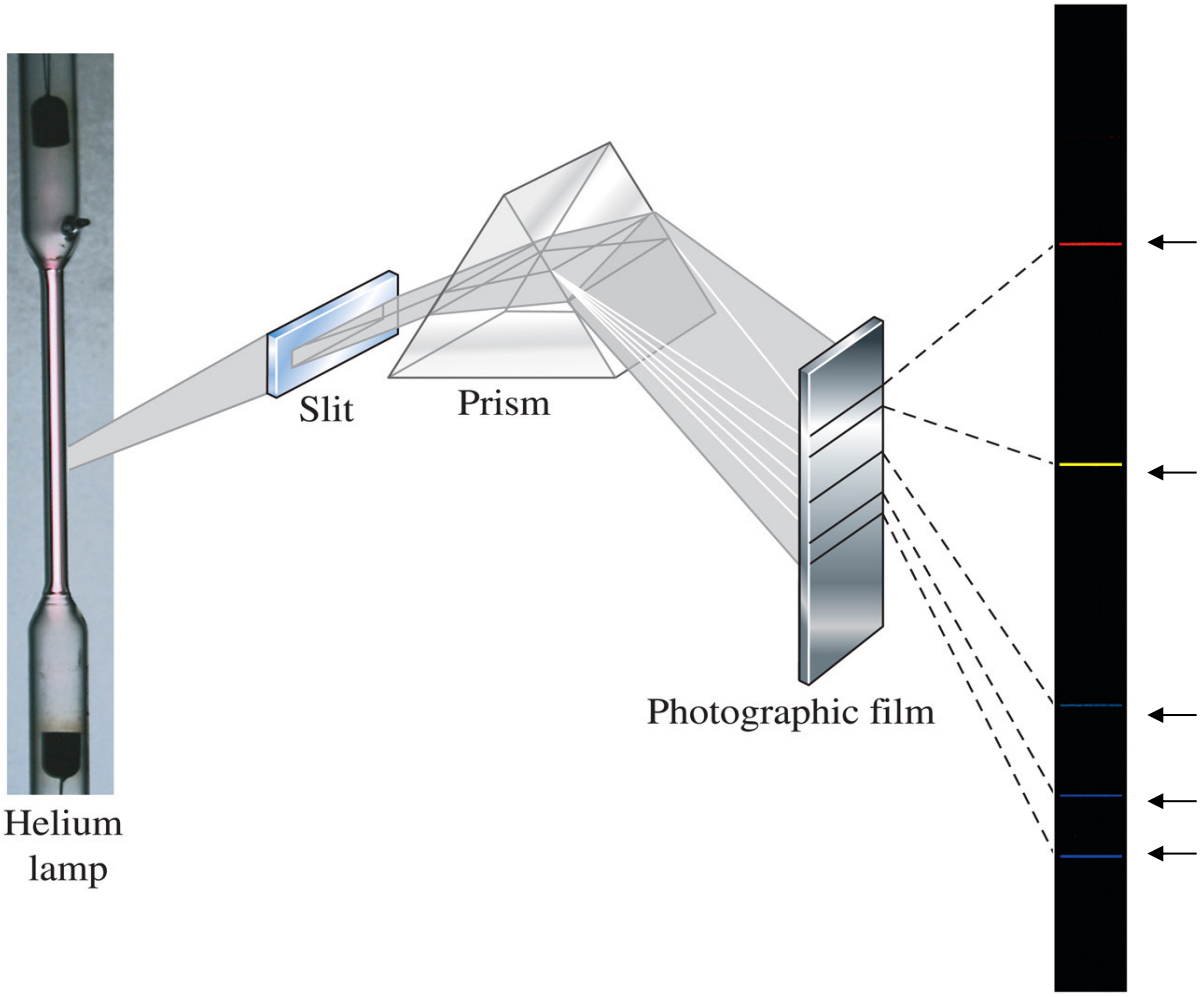


▲
Licht-Brechung

(a) (b) (c) (d) (e)

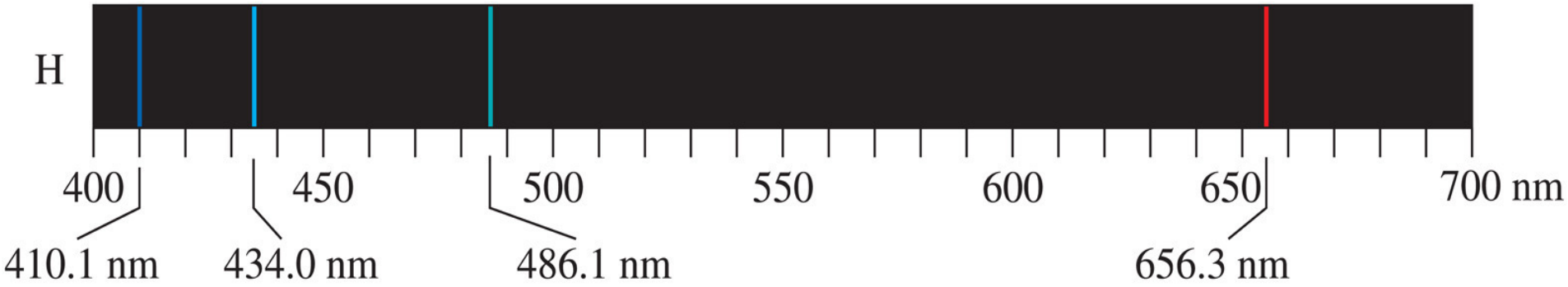
▲

Quellen für Lichtemission



▲

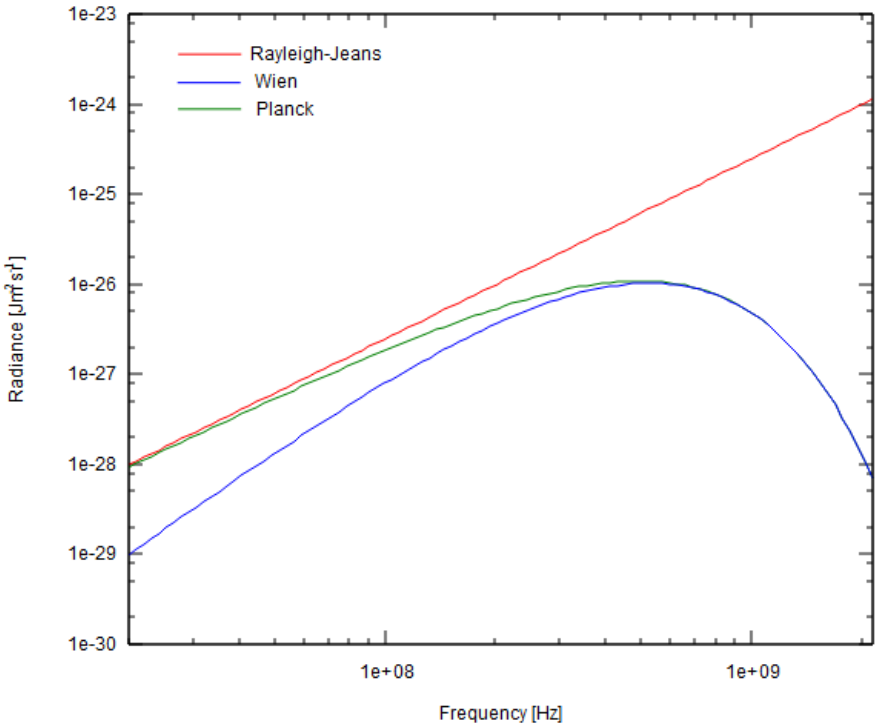
Linienspektrum von Helium

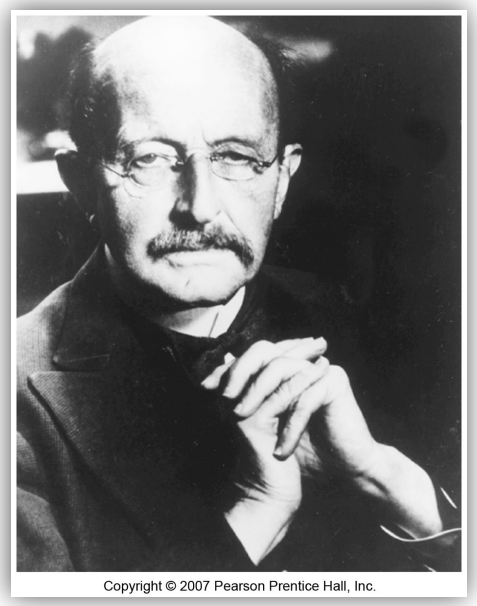


▲

Balmer-Serie für Wasserstoff-Atome - Linienspektrum

Quantentheorie





Max Planck, 1900

Energie ist wie Materie

diskontinuierlich

$$E = h\nu$$

▲

Strahlungspektrum eines erhitzten Gegenstandes

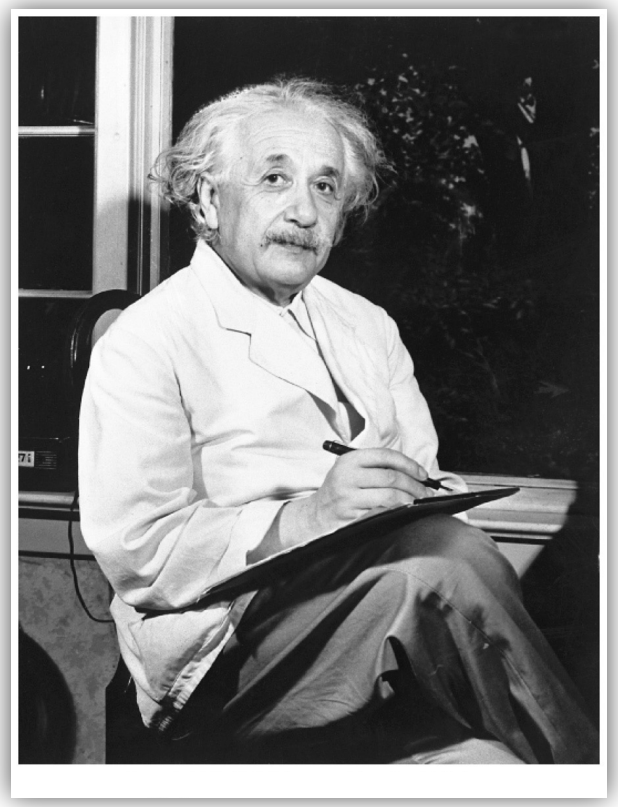
▲ **Lichtemission von geschmolzenem Eisen**

Der Photoelektrische Effekt

Heinrich Hertz, 1888

Licht, das auf eine metallische Oberfläche trifft, führt zur Freisetzung von Elektronen.

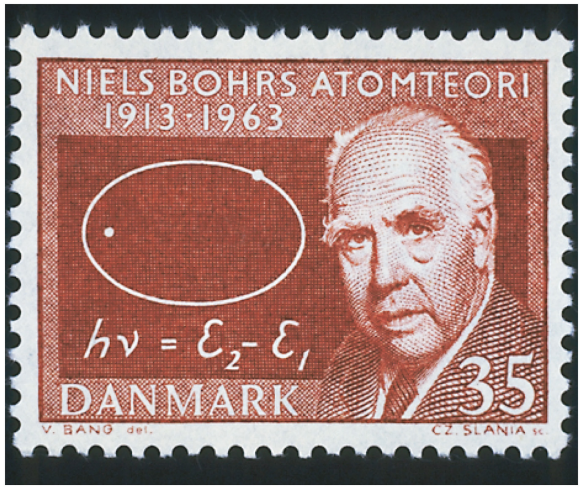
$\nu > \nu_o$

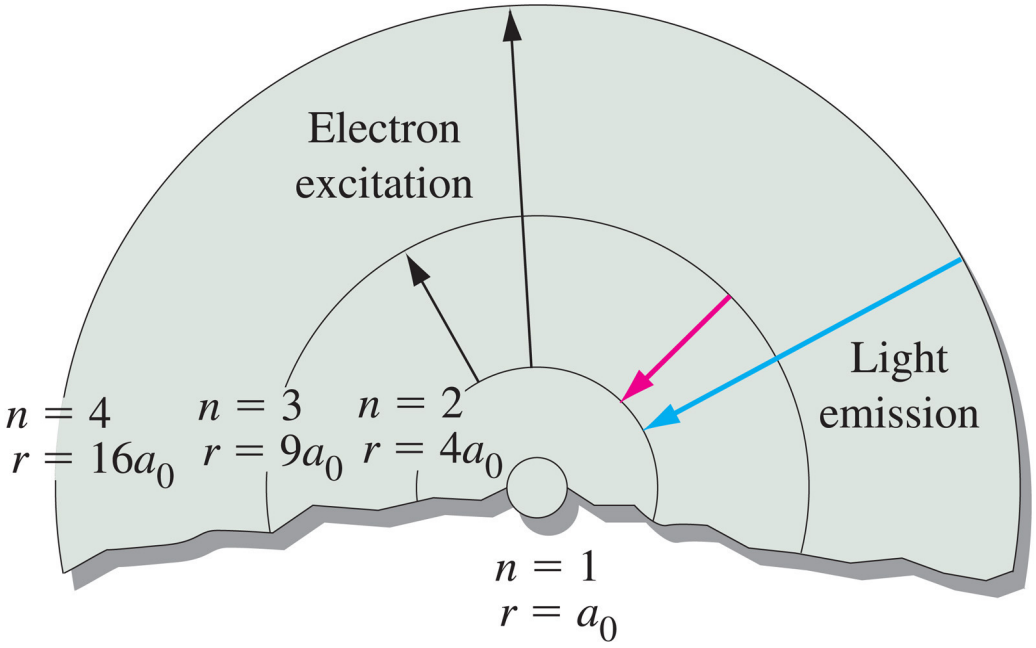


Albert Einstein 1905

Das Bohr-Atommodell

$$E = \frac{-R_H}{n^2}$$

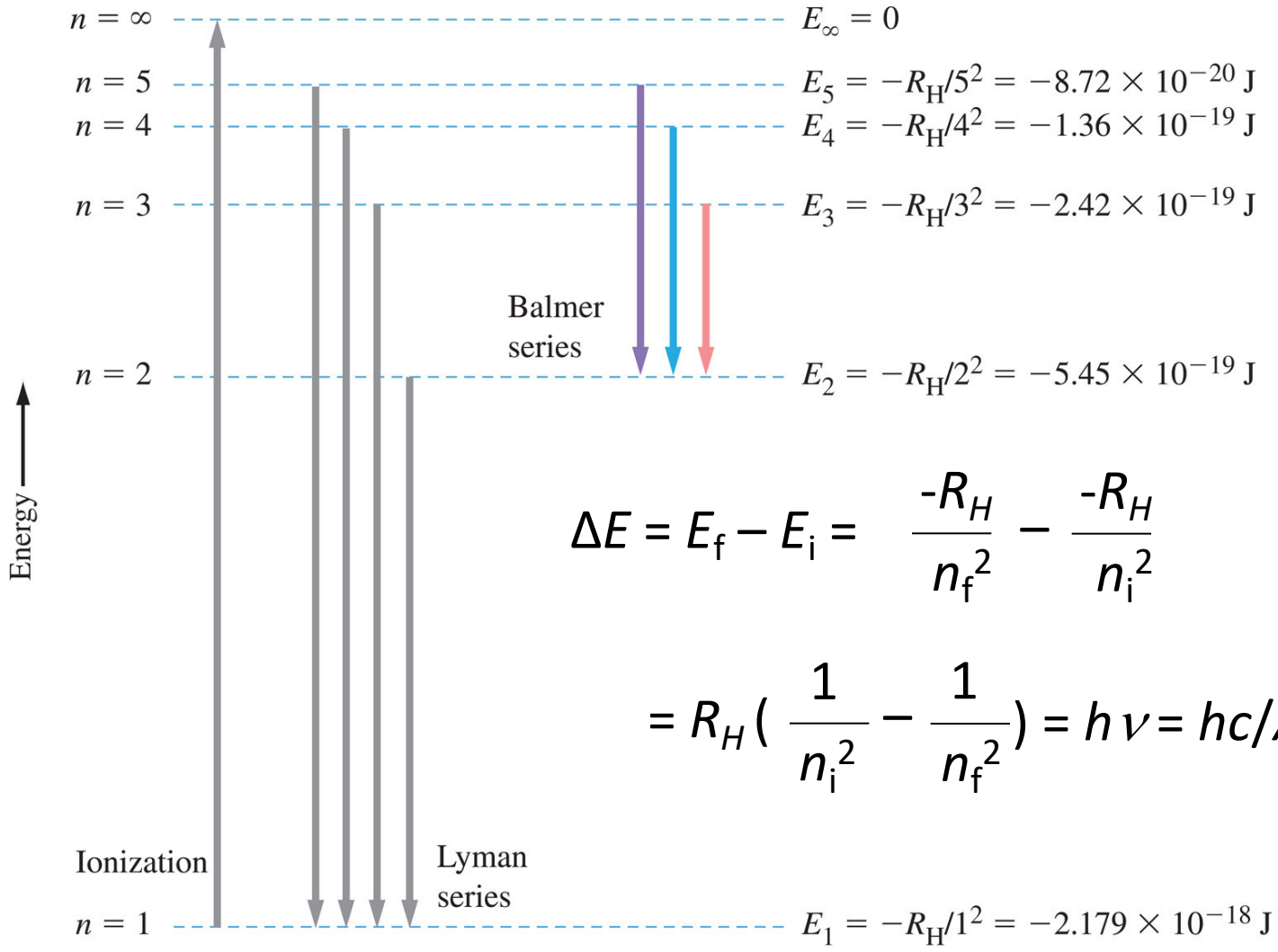




$R_H = 2.179 \times 10^{-18}$ J

▲

Bohr-Modell des Wasserstoff-Atoms

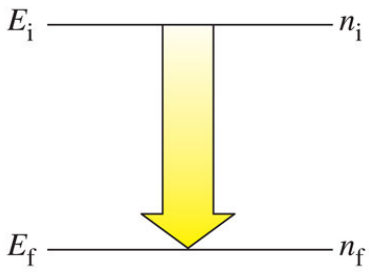


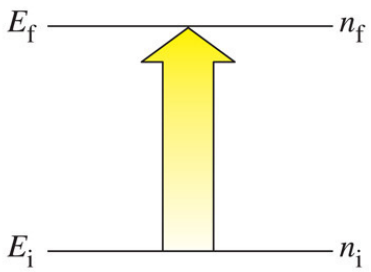
$$\Delta E = E_f - E_i = \frac{-R_H}{n_f^2} - \frac{-R_H}{n_i^2}$$

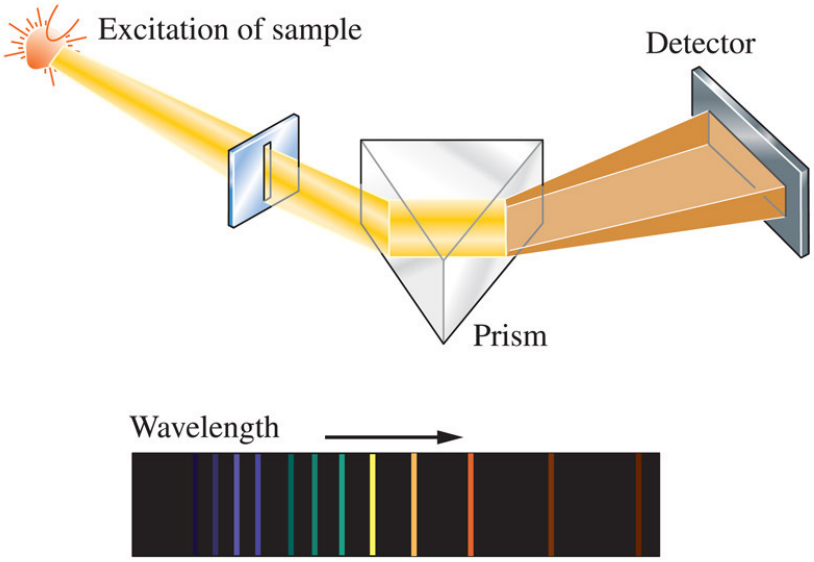
$$= R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = h\nu = hc/\lambda$$

▲

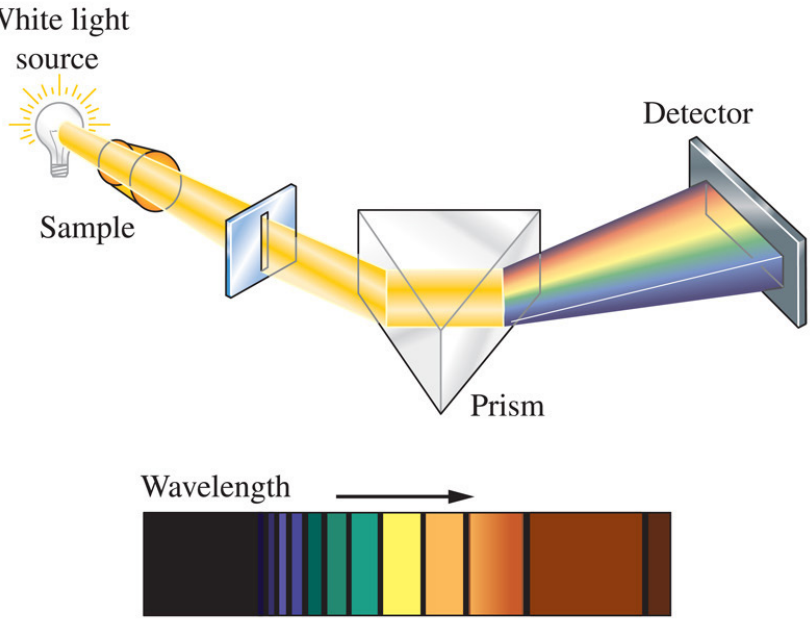
Energie-Diagramm für das Wasserstoff-Atom







Emission



Absorption

▲

Emissions- und Absorptions-Spektrum

Zwei Ideen die zur Quantenmechanik führten

Wellen-Teilchen-Dualismus

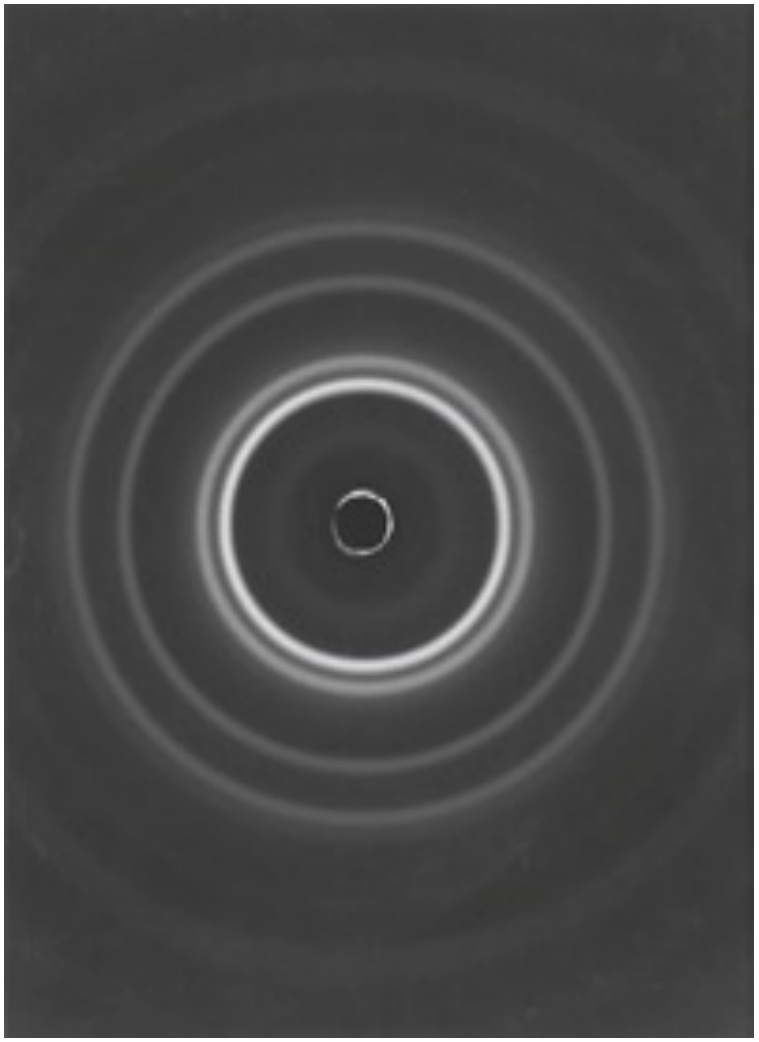
Einstein schlug vor, dass Teilchen-Eigenschaften den photoelektrischen Effekt erklären könnten.

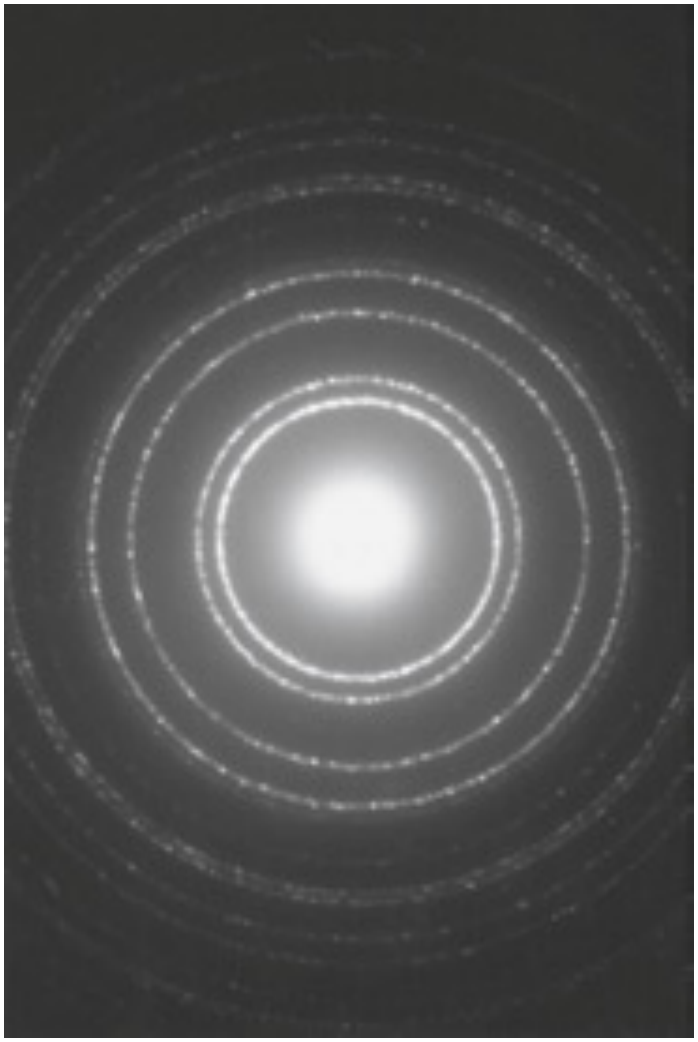
Beugungsmuster deuten auf wellenähnlichen Charakter von Photonen hin.

deBroglie, 1924

Kleine Teilchen können wellenähnliche Eigenschaften zeigen.

Louis de Broglie
Nobelpreis 1918

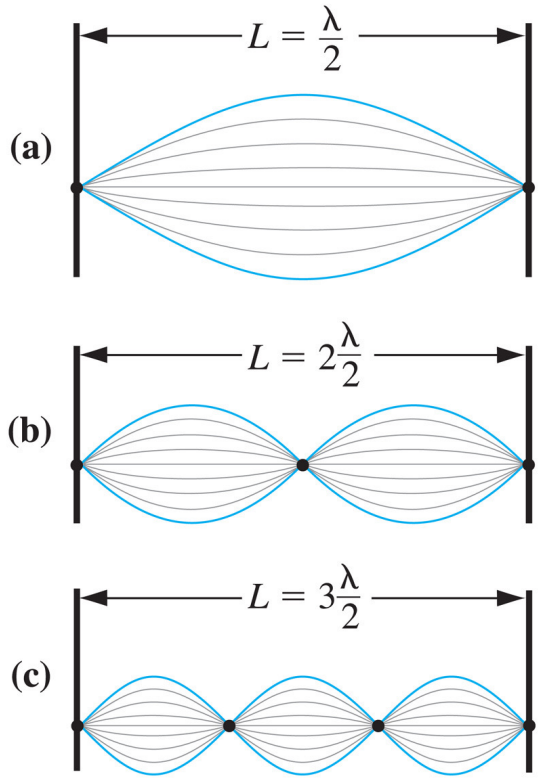




▲

Welleneigenschaften von Elektronen experimentell

Wellenmechanik



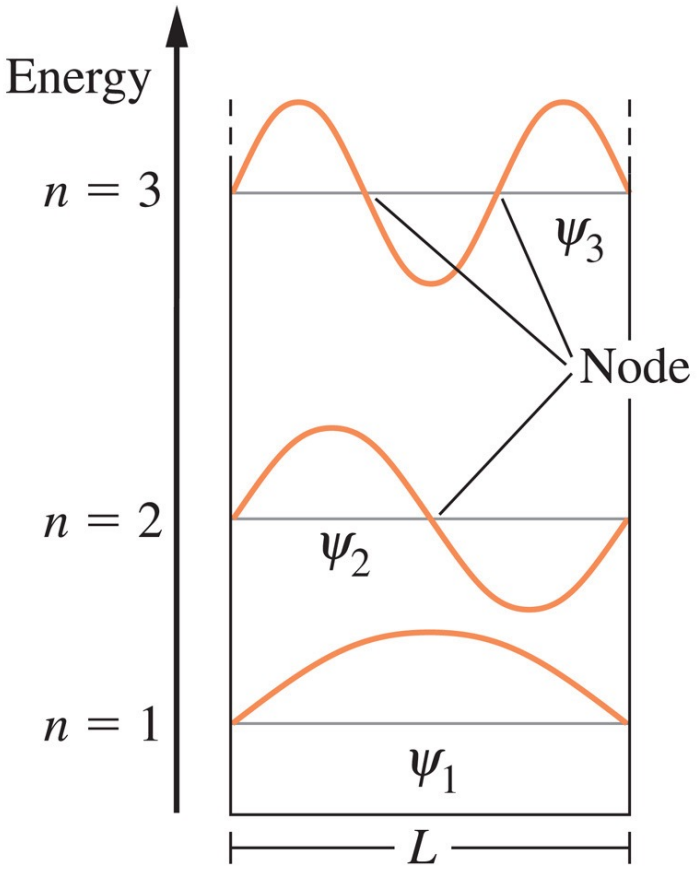
▲ **Stehende Wellen**

Stehende Wellen

Knoten werden nicht verschoben.

$$\lambda = \frac{2L}{n},\ n = 1, 2, 3\ldots$$

Teilchen in einer Box: Stehende Wellen, Quantenteilchen und Wellenfunktionen



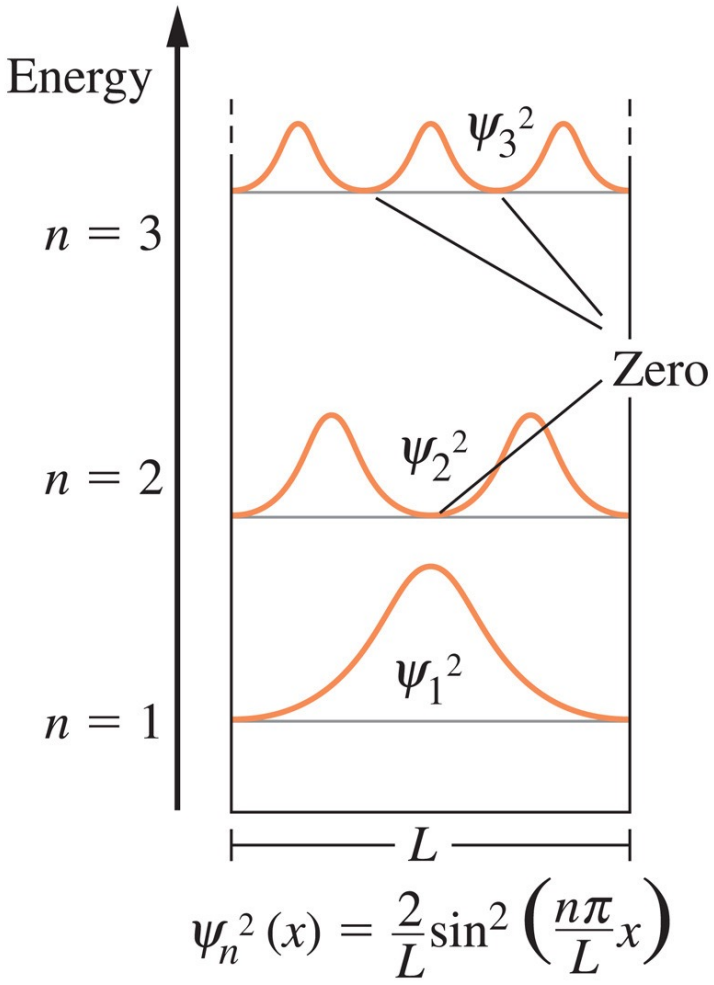
ψ, psi, die Wellenfunktion.

Teilchen in einer Box.

$$\psi = \sqrt{\frac{2}{L}} \sin\left(\frac{n\pi}{L} x\right)$$

▲

Stehende Wellen eines Teilchens in einer Box.



▲

Wahrscheinlichkeiten

Quantenzahlen und Elektronenzahl

Hauptquantenzahl, n = 1, 2, 3…

Nebenquantenzahl,

ℓ = 0, s

ℓ = 1, p

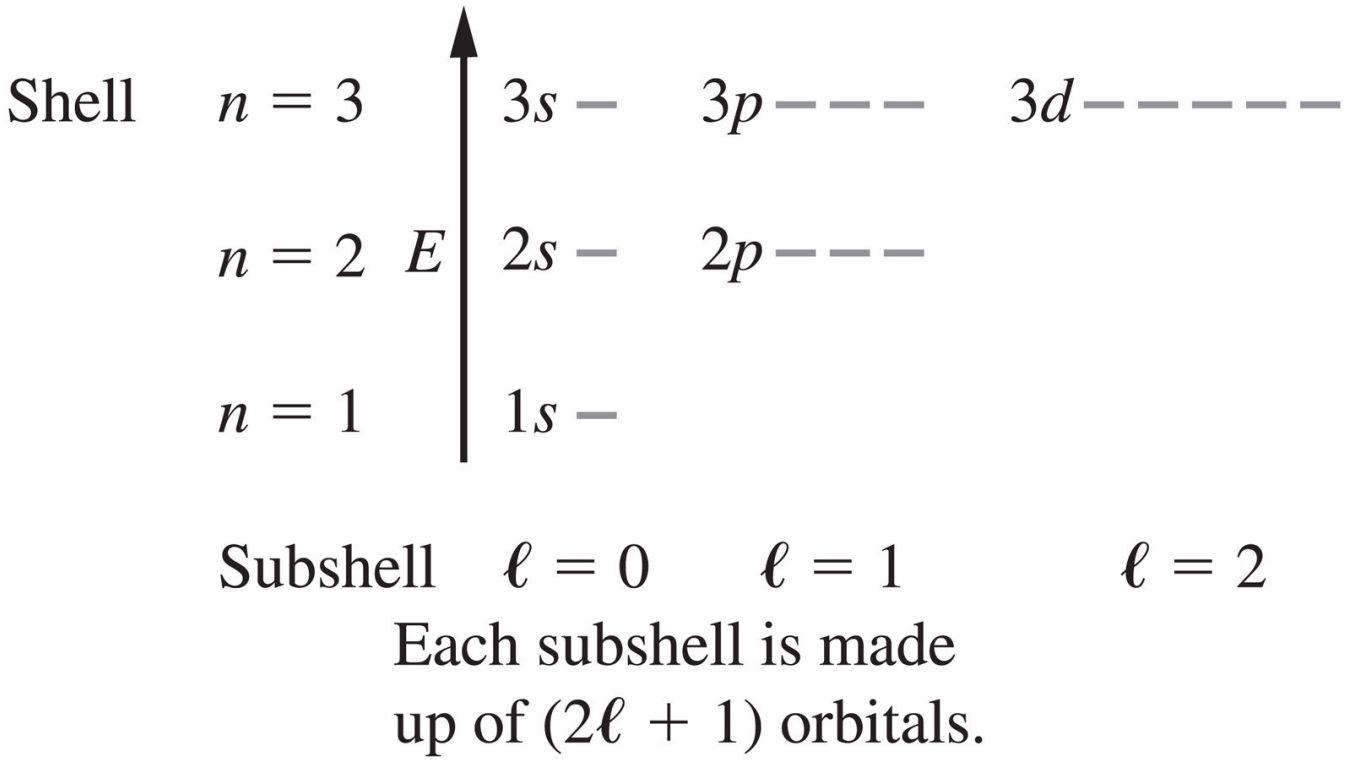
ℓ = 2, d

ℓ = 3, f

Magnetquantenzahl,

m_l = -3, -2, -1, 0, 1, 2, 3

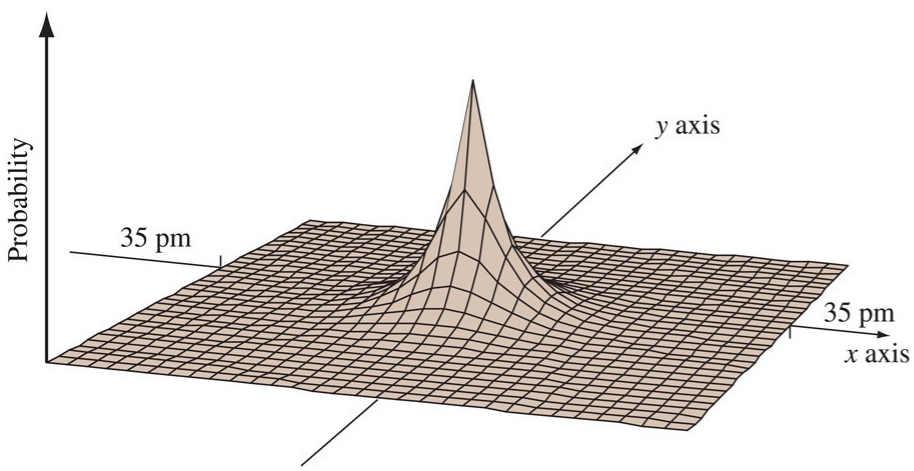
Haupt- und Unterenergieniveaus

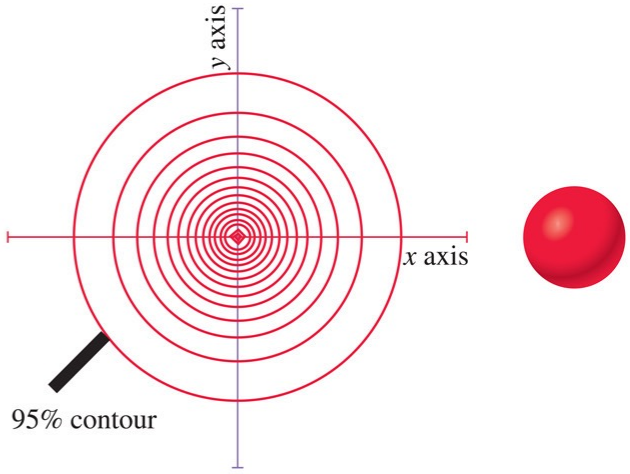


▲

Schalen und Unterschalen eines Wasserstoff-Atoms

s Orbitale

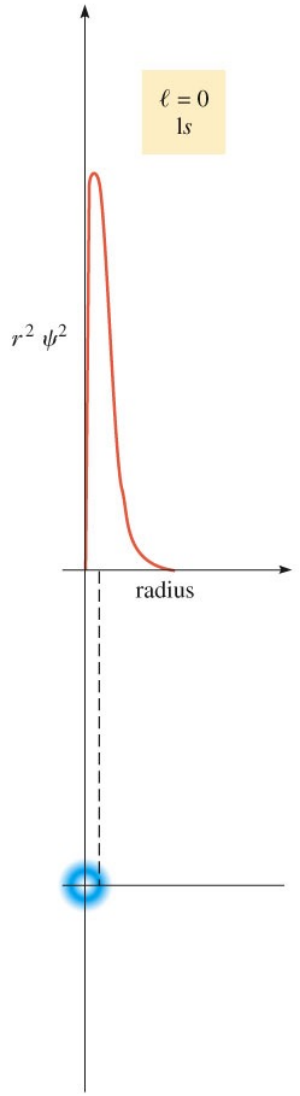




▲

3 Repräsentation für die Elektronen-Wahrscheinlichkeits-Dichten für das 1s-Orbital

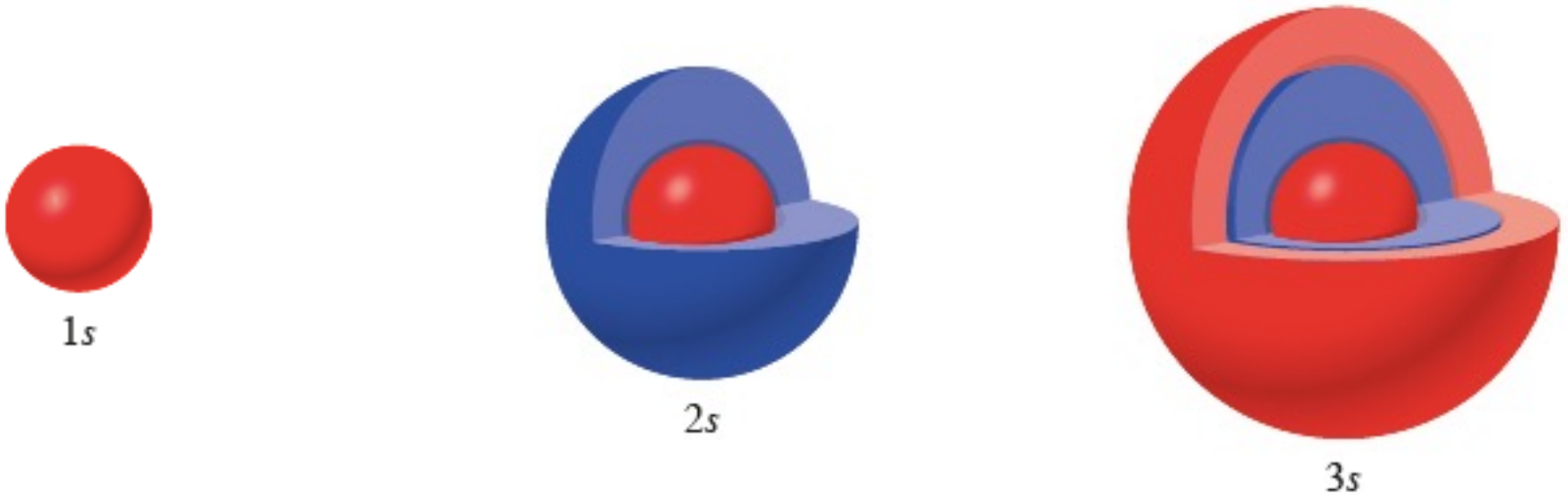
s Orbitale



▲

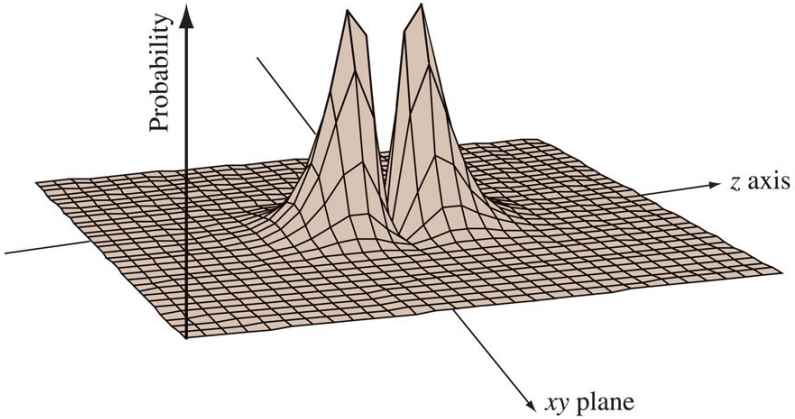
1s, 2s und 3s mit Querschnitten

2s Orbitale

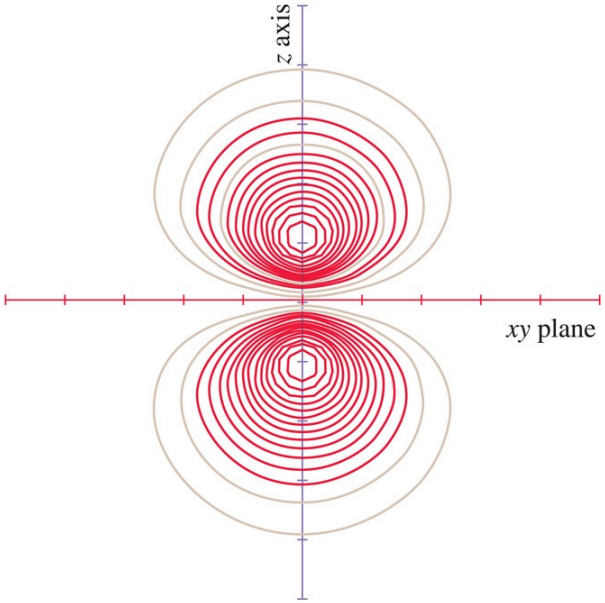


▲

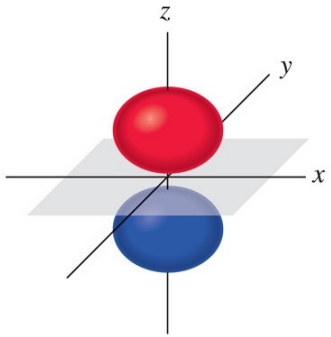
95% Elektronenwahrscheinlichkeitsdichten für 1s, 2s und 3s Orbitale



(a)



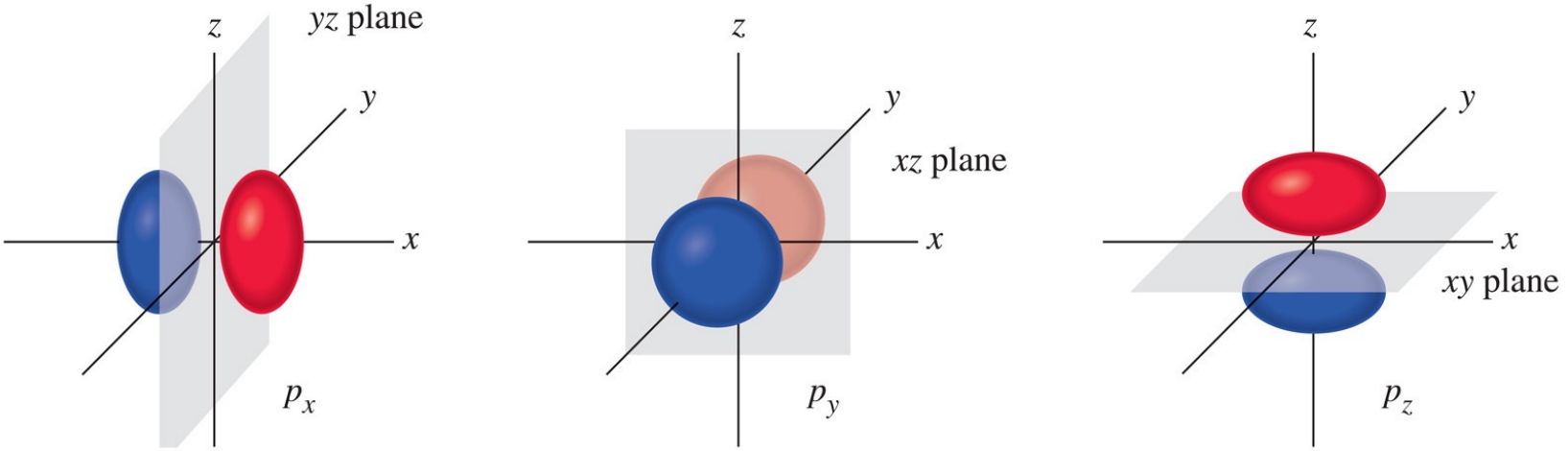
(b)



(c)

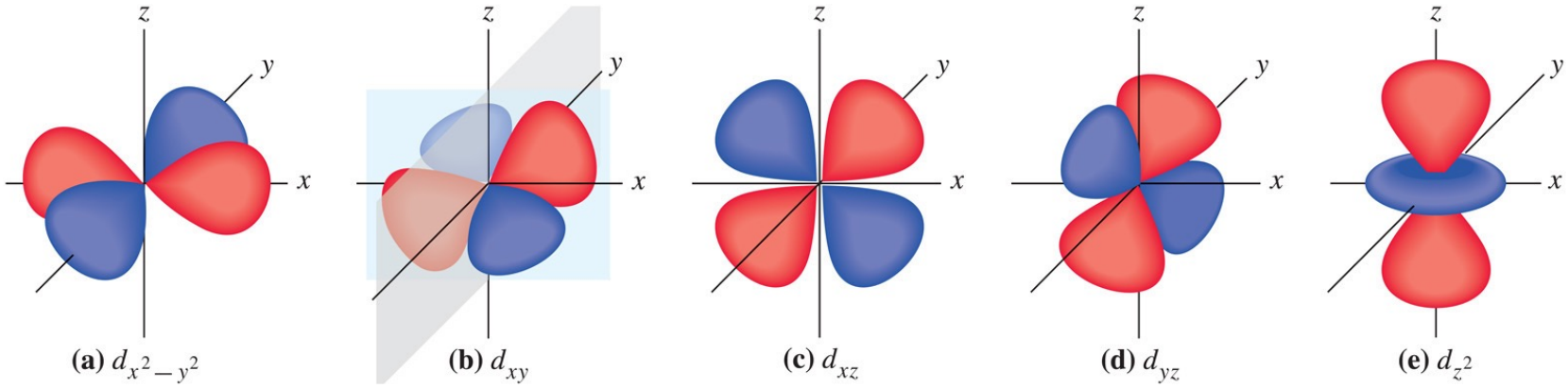
▲

2p Orbital



▲

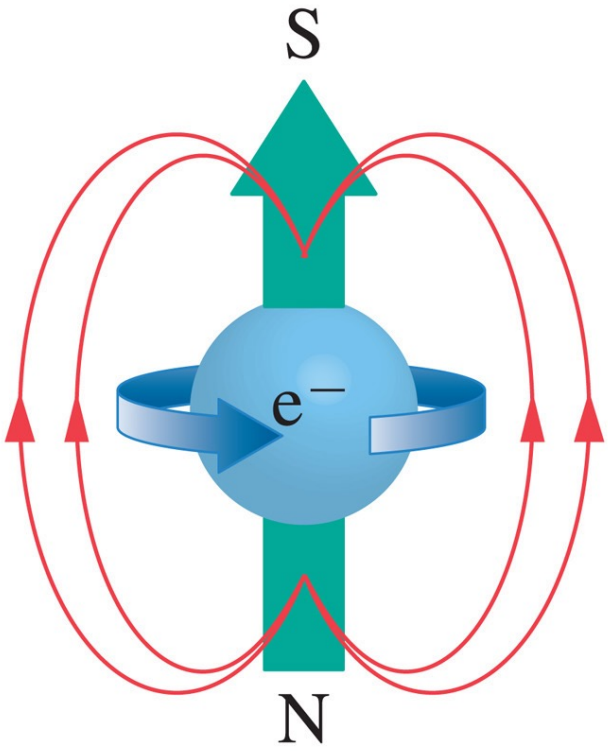
Die drei 2p-Orbitale



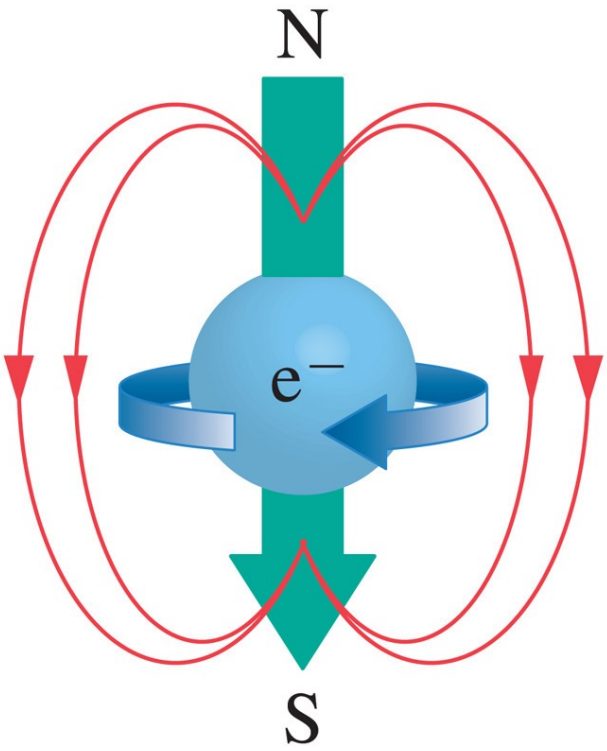
▲

Fünf d-Orbitale

Elektronenspin: Eine vierte Quantenzahl



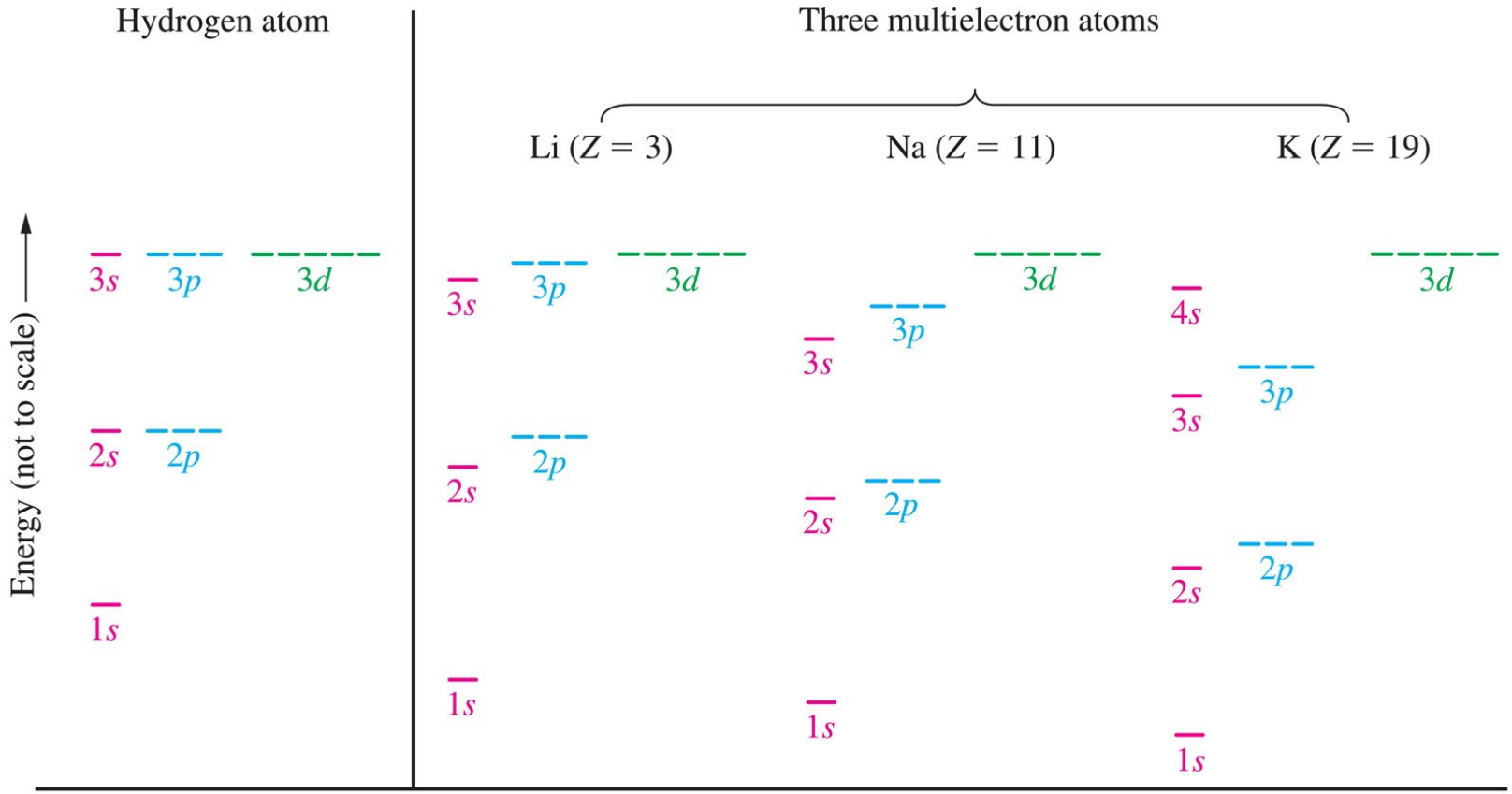
$$m_s = +\frac{1}{2}$$



$$m_s = -\frac{1}{2}$$

▲

Elektronenspin



▲

Energiediagramm für die ersten drei Schalen.

Elektronenkonfigurationen

Aufbau-Prinzip

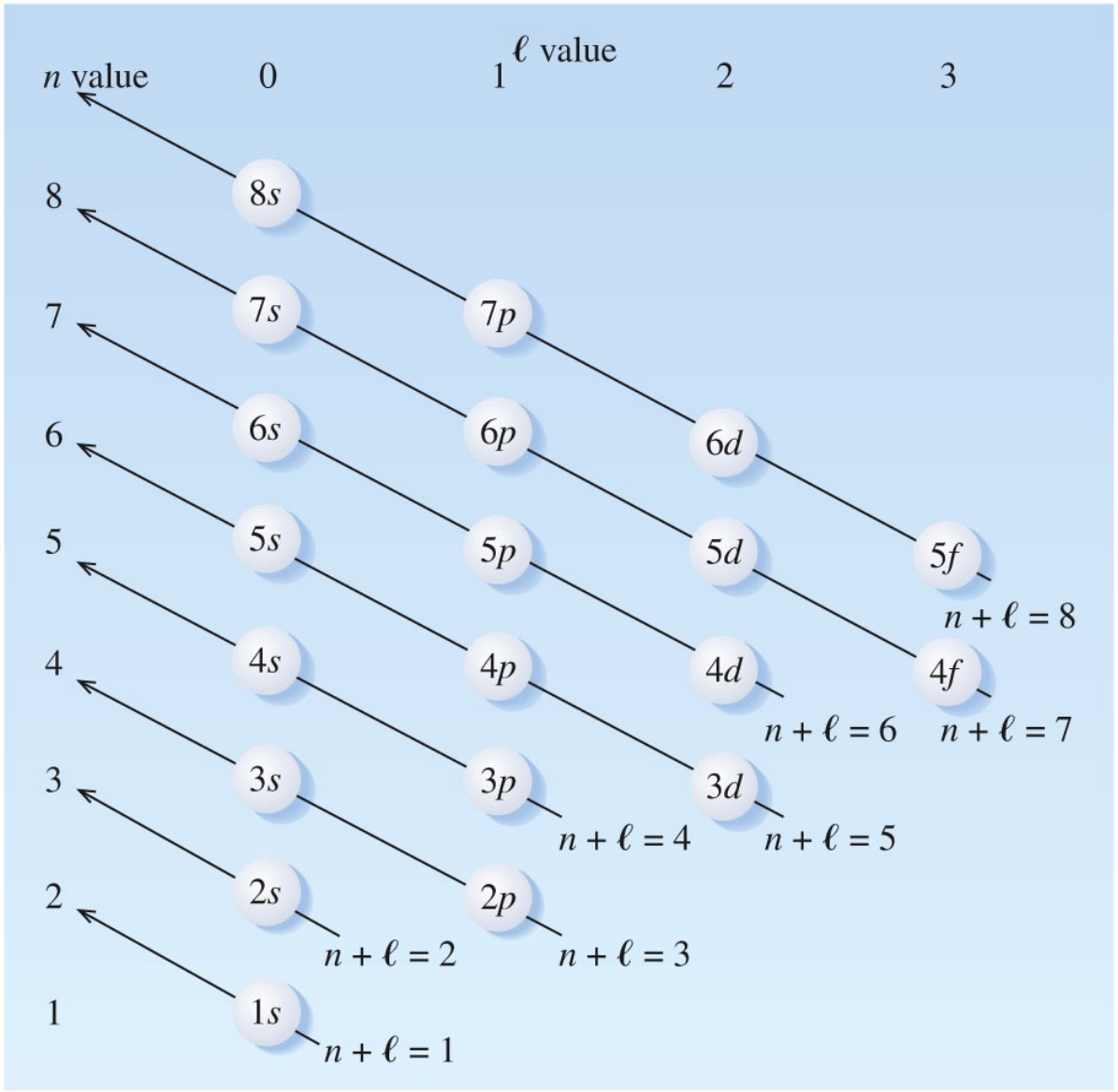
Von unten nach oben auffüllen.

Pauli-Prinzip

Elektronen müssen sich in mindestens einer Quantenzahl unterscheiden.

Hundsche Regel

Orbitale zuerst einzeln mit parallelen Spins besetzen.

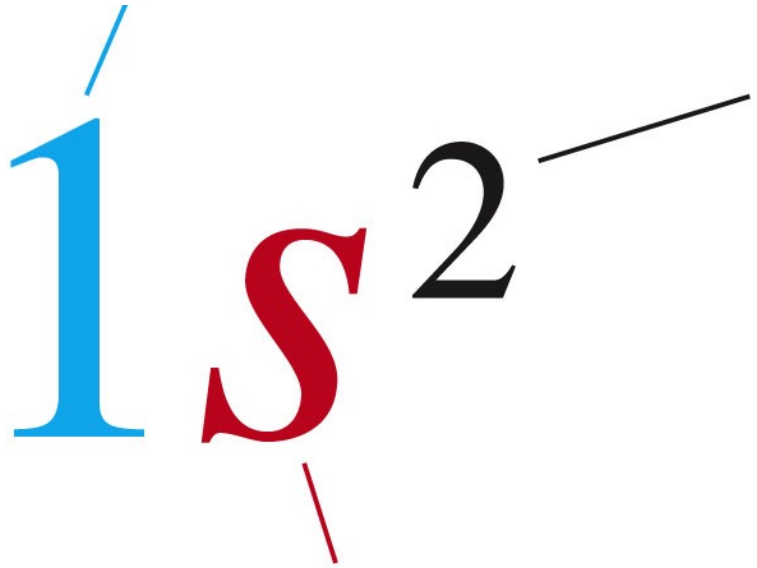


▲

Auffüllen von Unterschalen

Notation von Elektronenkonfigurationen

Hauptquantenzahl n



Anzahl Elektronen im Orbital

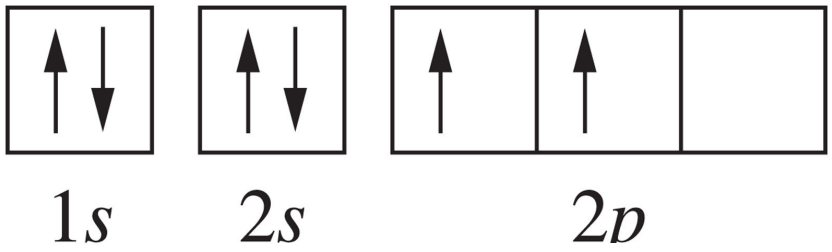
Nebenquantenzahl l

Notation von Elektronenkonfigurationen

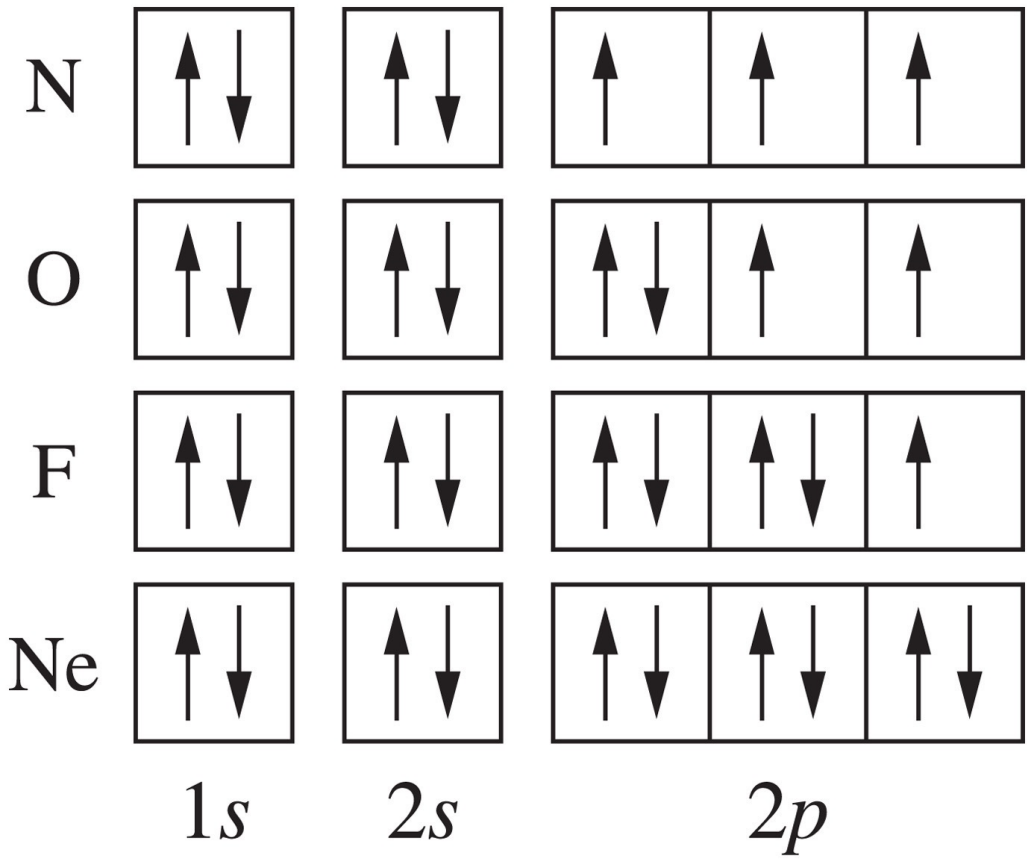
spdf Notation $1s^2 2s^2 2p^2$

spdf Notation (erweitert) $1s^2 2s^2 2p_x^1 2p_y^1$

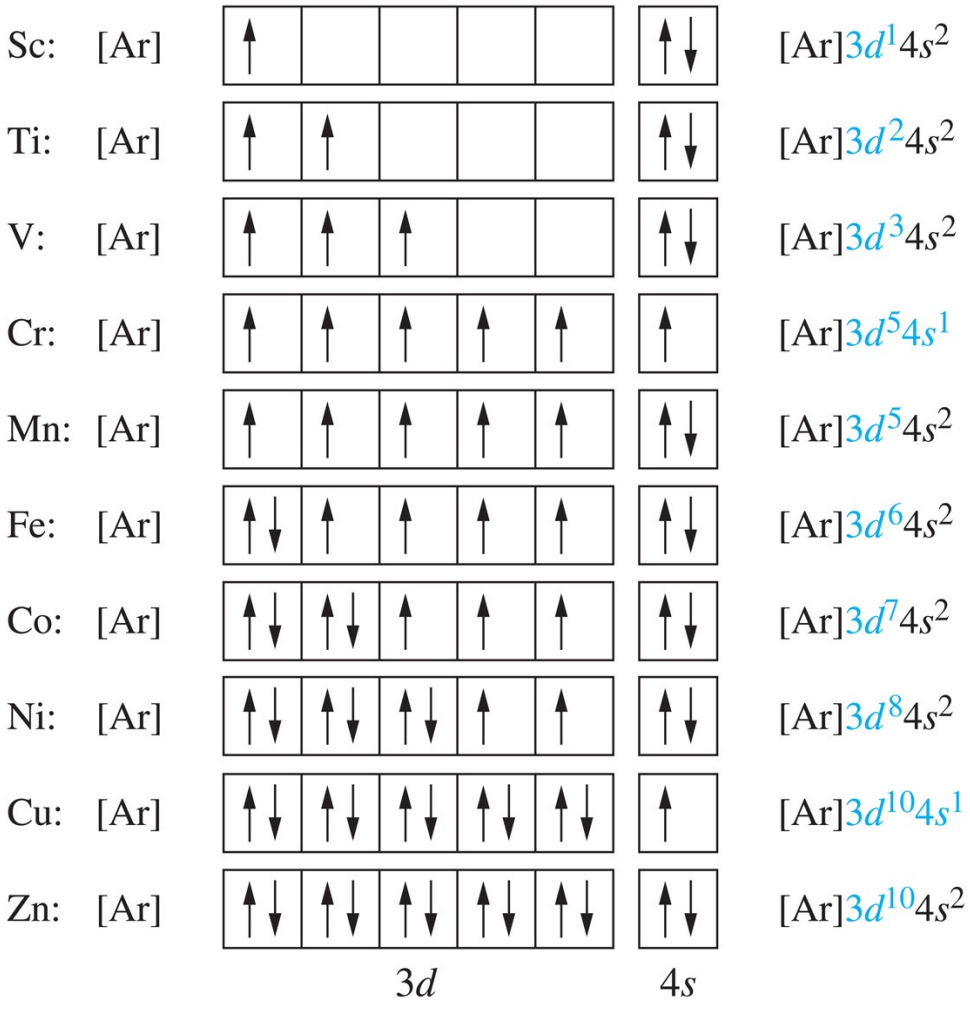
spdf Notation (Kästchen)



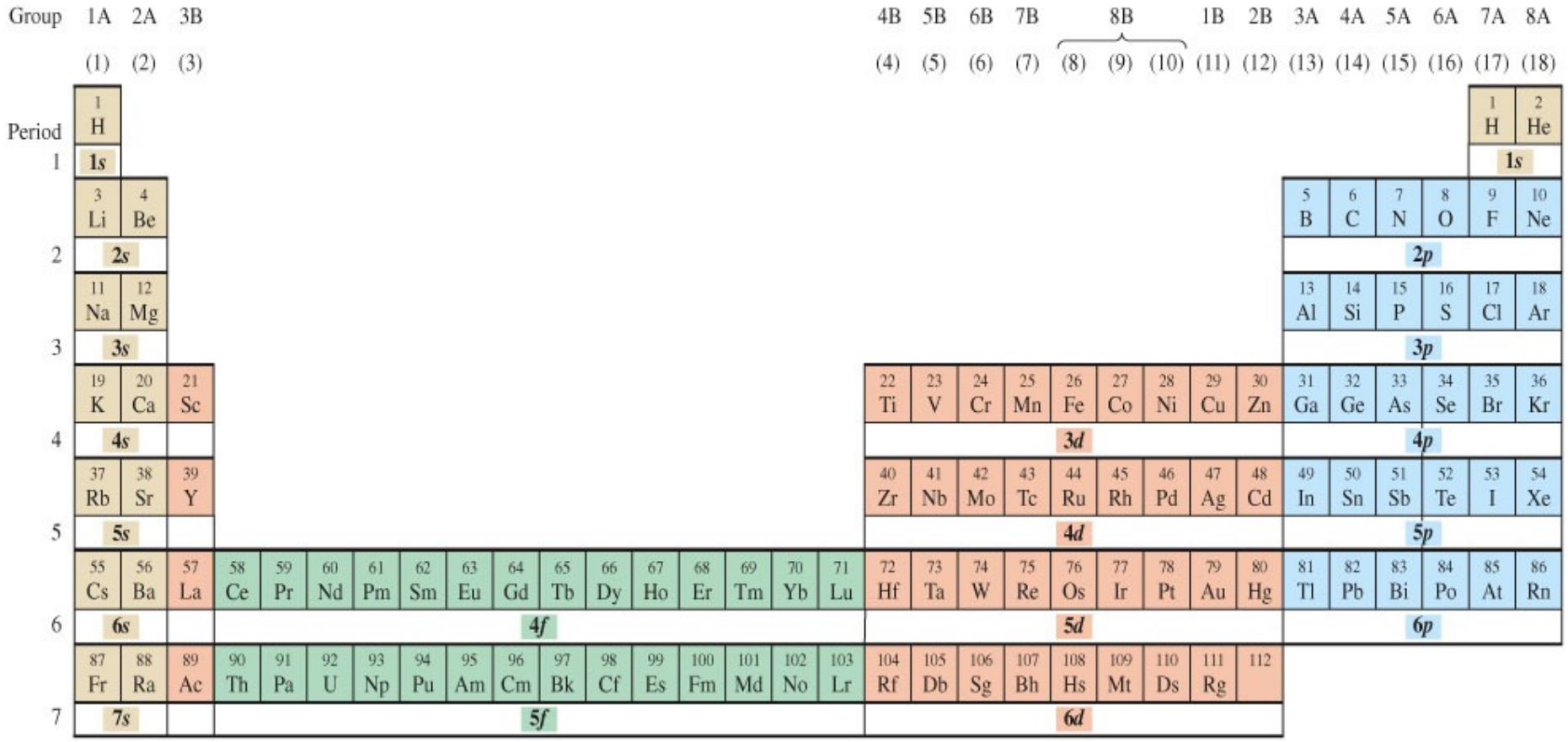
Aufbau-Prinzip



Aufbauprinzip – Sc bis Zn



Elektronenkonfiguration und das PSE



▲

Orbital-Blöcke

TABLE 8.2 Electron Configurations of Some Groups of Elements

Group	Element	Configuration
1	H	$1s^1$
	Li	$[He]2s^1$
	Na	$[Ne]3s^1$
	K	$[Ar]4s^1$
	Rb	$[Kr]5s^1$
	Cs	$[Xe]6s^1$
	Fr	$[Rn]7s^1$
17	F	$[He]2s^2 2p^5$
	Cl	$[Ne]3s^2 3p^5$
	Br	$[Ar]3d^{10} 4s^2 4p^5$
	I	$[Kr]4d^{10} 5s^2 5p^5$
	At	$[Xe]4f^{14} 5d^{10} 6s^2 6p^5$
18	He	$1s^2$
	Ne	$[He]2s^2 2p^6$
	Ar	$[Ne]3s^2 3p^6$
	Kr	$[Ar]3d^{10} 4s^2 4p^6$
	Xe	$[Kr]4d^{10} 5s^2 5p^6$
	Rn	$[Xe]4f^{14} 5d^{10} 6s^2 6p^6$