Cathode-Ray Tubes

When electricity is applied to a sealed glass tube containing a gas at low pressure the tubes appears to glow.

1870s, William Crookes observed that cathode rays are attracted by a magnetic field.

1897, Thomson demonstrated that cathode rays are deflected by an electric field as well.

1886, Goldstein discovered positive rays using cathode-ray tubes with small hold in the cathode.

Thomson Model of the Atom

Rutherford Model of the Atom

Atomic Mass of an Element

The atomic mass of an element is the weighted average mass of all naturally occurring isotopes.

Isotope	Mass	Abundance
^{12}C	12.000 amu	98.89%
^{13}C	13.003 amu	1.11%

^{12}C 12.000 amu x 0.9889 = 11.87 amu

^{13}C 13.003 amu x 0.0111 = 0.144 amu

12.01 amu

The Quantum Concept

1900, Max Planck proposed that radiant energy is not continuous, but rather, the radiation is emitted in small bundles.

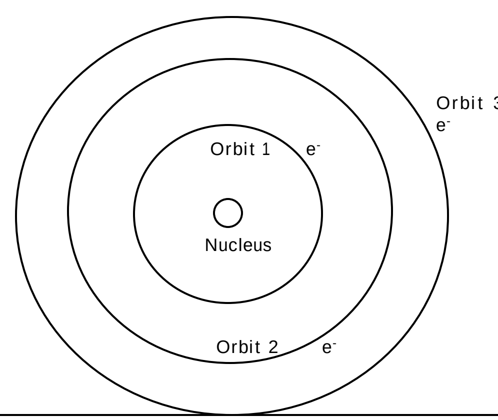
The idea that energy is released in discrete unit is referred to as the quantum concept.

The Bohr Model of the Atom

1913, Bohr speculated that electrons orbit around the atomic nucleus just as planets circle around the Sun.

He suggested that the electron orbits were at a fixed distance from the nucleus and had a definite energy.

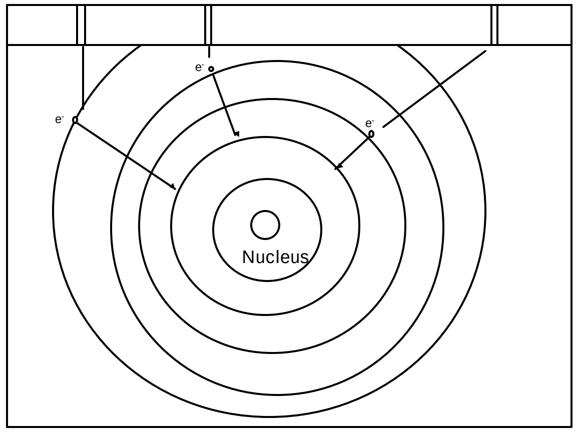
The electron was said to travel in a fixed-energy orbit that was referred to as an energy level.



Evidence for Electrons in Fixed-Energy Levels

The collection of narrow bands of light energy is referred to as an **emission line spectrum**, and the individual bands of light are called spectral lines.

The concept of electron energy levels is supported by spectral lines.



Atomic Fingerprints

The study of emission spectra revealed that each element produced a unique set of spectral lines.

This observation indicated that the energy levels must be unique for atoms of each element. A line spectrum is used as an "atomic fingerprint."

The model proposed by Niels Bohr was supported experimentally by the emission spectrum of hydrogen.

The emission spectrum of other elements besides hydrogen had far too many lines to interpret.

The model that eventually emerged had electrons occupying a **energy sub-level** with a main energy level. These energy sub-levels were designated s, p, d, and f in reference to the *sharp*, *principal, diffuse,* and *fine* lines in the emission spectra of the elements.

S sublevel = 2 e^-

p sublevel = 6 e^-

d sublevel = 10 e^-

Energy level	Energy Sublevel	Maximum # e^- in Sublevel	Maximum # e^- in Energy level
1	1 s	2 e^-	2 e^-
2	2 s	2 e^-	
	2 p	6 e^-	8 e^-
3	3 s	2 e^-	
	3 p	6 e^-	
	3 d	10 e^-	18 e^-
4	4 s	2 e^-	
	4 p	6 e^-	
	4 d	10 e^-	
	4 f	14 e^-	32 e^-

Order of electron filling

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s

The electron configuration of an atom is a shorthand statement for describing the location of the electrons by sublevel. First, the sublevel is written, followed by a superscript that indicates the number of electrons.

Ne: $1s^2$ $2s^2$ $2p^6$

Cl: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$

C: $1s^2$ $2s^2$ $2p^2$

Cu: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^9$

In 1920 Werner Heisenberg suggested the uncertainty principle; that is, it is impossible to simultaneously know the precise location and energy of an electron. Instead, the energy of an electron can only be known in terms of its probability of being located somewhere within the atom. This description gave rise to the quantum mechanical atom. A location within the atom where there is a high probability of finding a electron having a certain energy is called an orbital.

The shape of an *s* orbital is spherical; a *p* orbital resembles the shape of a dumbbell.