Advanced Chemistry
Notes: Atomic Theory I

Democritus (470-400 BC)
John Dalton (1766-1844)

Dalton's Atomic Theory (1808)

1. An element is composed of extremely small, indivisible particles called atoms.
2. All atoms of a given element have identical properties that differ from those of other elements.
3. Atoms cannot be created, destroyed, or transformed into atoms of another element.
4. Compounds are formed when atoms of different elements combine with one another in small whole-number ratios.
5. The relative numbers and kinds of atoms are constant in a given compound.

Humphrey Davy (1778 – 1829)

Observation: Electrolysis – pass electric current through some substances they will decompose
Conclusion: chemical compounds are held together by electrical forces

Cathode Ray Tube Studies

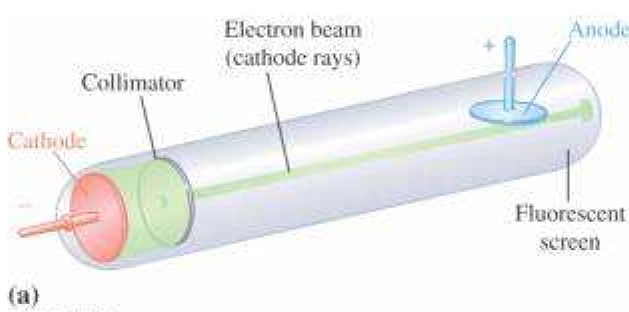
Supports hypothesis of existence of electrons (negatively charged particles associated with atoms)

CRT:

- gas at very low pressure
- high voltage applied across tube
- electrons pulled from the cathode to the anode (a current flows)

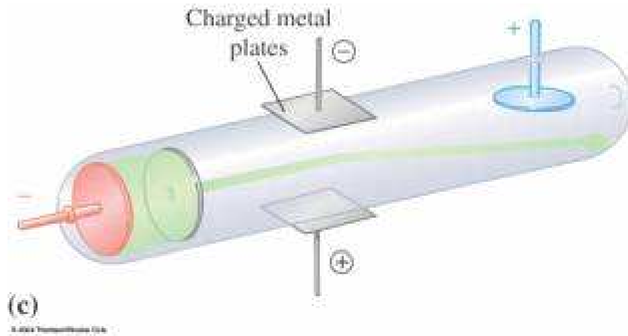
Experiments

- **Observation:** A screen placed at the anode glows when struck by the e -
- **Conclusion:** particles are traveling from the cathode to the anode

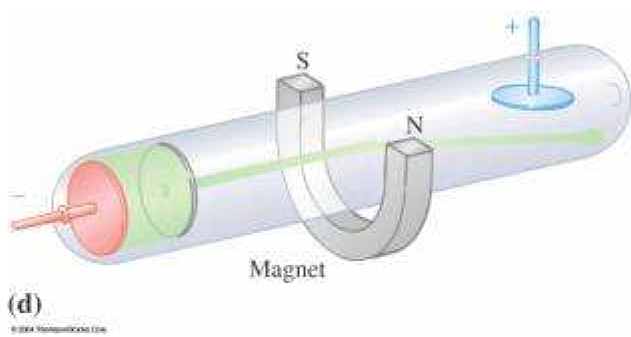


- **Observation:** When a tiny object is placed in the middle of the tube, a shadow is cast on the screen at the anode
- **Conclusion:** the particles travel in a straight line

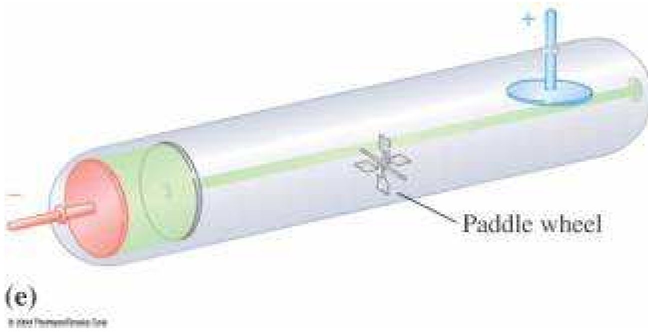
- **Observation:** The cathode rays are deflected by an electric field
- **Conclusion:** The particles are negatively charged



- **Observation:** The cathode rays are deflected by a magnetic field
- **Conclusion:** The particles are moving negative charges



- **Observation:** A paddle wheel in the middle of a cathode ray tube turns
- **Conclusion:** The particles have mass



J. J. Thomson (1856 – 1940)
1897, Credited with the discovery of the electron

Observations:

- Using a cathode ray tube quantified above studies

- Found the **ratio of the charge to the mass of the particles in the cathode rays** using the laws of physics and equations relating force, mass, electric field strength and magnetic field strength.

- The charge to mass ratio is always the same regardless of the gas used, the voltage, etc.

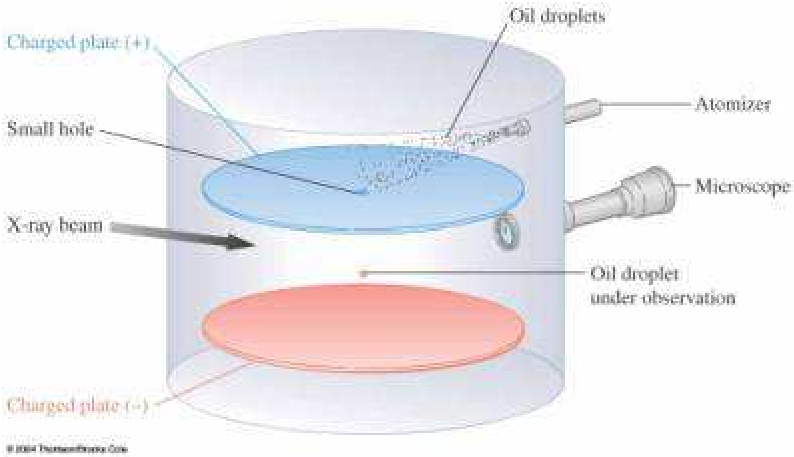
Conclusion:

- Since the charge to mass ratio was always the same – he determined that these negative particles are **fundamental particles of matter (electrons)**

Robert Millikan (1868 – 1953)

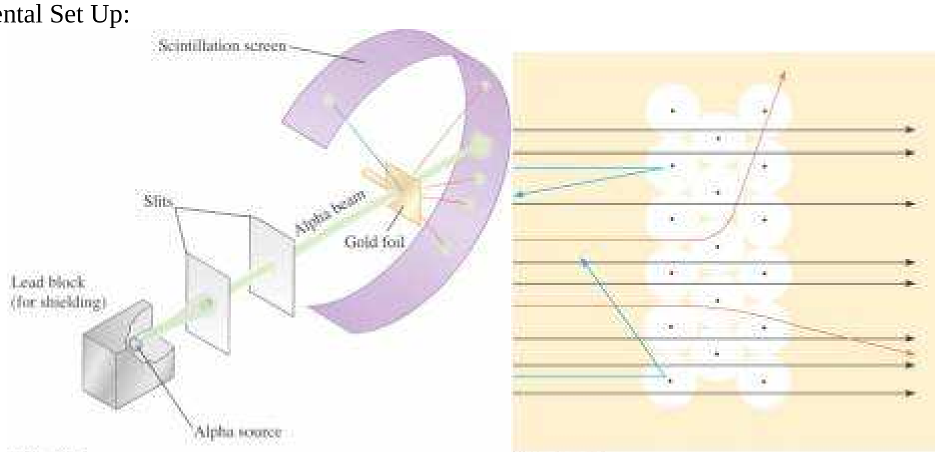
1909, Millikan Oil Drop Experiment
Experimental Set Up:

- Small drops of oil with a negative charge are examined.
- The diameter of a drop is measured. From this the volume and subsequently the mass are calculated (using the known density of the oil).
- By adjusting the electric field to balance the force of gravity, the charge on the drop can be calculated.
- The charge of millions of drops was calculated.



Observation: The charge of each drop is a whole number multiple of some common number.
Conclusion: Note in the data above, each number is divisible by 1.92 x 10 -19 ; this is the charge of an electron.

Ernest Rutherford (1871 – 1937)
1911, Discovery of the nucleus
Experimental Set Up:



- α Particles were known to be heavy positive particles.
- When the particles hit the screen brief flashes of light appear

- Rutherford's experiment is designed to determined to arrangement of the positive and negative particles in the atom based on the plum pudding model
- A beam α particles should go right through the diffuse plum pudding with small deflections as it hit the diffuse positive charge

Observation: Most particles went straight through the foil un-deflected
Conclusion: The atom is mostly empty space
Observation: Some particles bounced back!
Conclusion: Most of the mass and all of the positive charge is centered in a small, dense central region which he called **the nucleus**

Rutherford Model of the Atom

- Atoms consist of very small, very dense positively charged nuclei
- The electrons are in clouds surrounding the nucleus at relatively large distances
- Mostly of the atom is empty space, and if the nucleus were the size of a dime, the closest electron would be a football field away.

James Chadwick (1891 – 1974)
1932 Identification of the neutron
Observation: When elements are bombarded with α particles, particles are emitted
Conclusion: These particles are **neutrons**, fundamental particles with no charge and a mass approximately equal to the proton.

Atomic Number (Z) = # protons in the nucleus

- This number is unique for each element
- Every gold atom has 79 protons, if a proton is taken away, you get Platinum
- For a neutral atom, the # protons = # electrons

Mass Number (A) = (# protons) + (# neutrons)

Notation for elements: $^{A}_{Z}E$

Example: for Carbon-13: $^{13}_{6}C$

Isotopes
Atoms of the same element must have the same number of protons, but they can have differing numbers of neutrons.

An isotope of an element is identified by the mass number

Carbon 12 has 6 protons, 6 neutrons
Carbon 13 has 6 protons, 7 neutrons