

Chemistry

Atomic Structure and the Periodic Table

Lesson 4

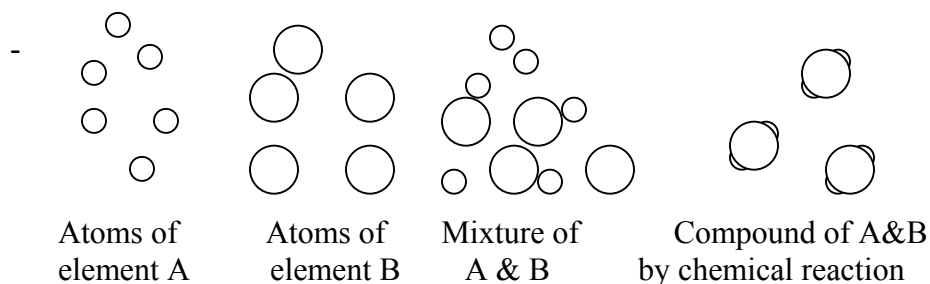
Lesson Plan

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Atoms

Objective: Summarize Dalton's atomic theory, Describe the size of an atom

- Atoms
 - Do you believe in things you cannot see?
 - Atom was first suggested in 4th century BC by Democritus
 - Indivisible and indestructible
 - John Dalton is credited with current beginning of understanding of the atom
 - Dalton's atomic theory
 - All elements are composed of tiny invisible particles called atoms
 - Atoms of the same element are identical. Atoms of any one element are different from those of any other element.
 - Atoms of different elements can physically mix together or chemically combine with one another in simple whole-number ratios to form compounds.
 - Chemical reactions occur when atoms are separated, joined or rearranged. Atoms of one element, however, are never changed into atoms of another element as a result of a chemical reaction.
 - Alchemy – transmutation



- Size of an atom
 - Imagine trying to grind up a penny – then cut up the fine grounds – you could never cut it down to an atom, blade too wide.
 - 100,000,000 atoms = 1 cm
 - 10^8 atoms = 1 cm
 - 1 penny has $\sim 2.4 \times 10^{22}$ atoms
 - The earth has 6×10^9 people
 - 1 penny contains 4×10^{12} more atoms than on the earth

- Structure of the Nuclear Atom
- *Objective: Distinguish among protons, electrons and neutrons in terms of relative mass and charge; Describe the structure of an atom, including the location of protons, electrons and neutrons with respect to the nucleus*
 - Change from Dalton's theory – atoms can be divided
 - There are 3 major sub-atomic particles – but there are dozens of kinds of sub-atomic particles
 - Electrons – negatively charged subatomic particles
 - JJ Thomson discovered electrons in 1897
 - Low pressure gas tube with a metal disk at each end
 - Metal disk called an electrode
 - Electrodes connected to high voltage
 - Anode became positively charged
 - Cathode became negatively charged
 - Glowing beam formed between the electrodes – traveling from the cathode to the anode – known as the cathode ray
 - Cathode rays would be attracted to metal plates with a positive charge, but repelled from plates with a negative charge.
 - Knowing that opposite charges attract and like charges repel, he proposed that the cathode ray was a stream of tiny negative charges moving at high speed .
 - He named them electrons
 - Proved that the type of gas or metal plates did not affect the production of electrons.
 - Concluded that electron was part of an atom
 - Mass of an electron is 1/2000 that of a hydrogen atom
 - Robert Millikan later determined the ratio of the charge to the mass of an electron
 - Electron is exactly 1 negative charge
 - Electron mass is 1/1840 of the mass of a hydrogen atom
- Protons & Neutrons
 - If a cathode ray is electrons – what remains of the atom?
 - Hydrogen – strip electron, what is left behind?
- 4 simple ideas to think through
 - Atoms have no charge (they are electrically neutral)
 - Electrical charges are carried by particles of matter
 - Electrical charges exists in whole numbers – no fractions
 - When a given number of negatively charged particles combine with an equal number of positively charged particles, an electrically neutral particle is formed.

- A positive particle was discovered in 1886 by Goldstein – observing a cathode ray and found a positively charge stream moving in the opposite direction as the cathode ray
 - Found to have charge of +1
 - Mass of 1840 times that of an electron
- 1932 James Chadwick confirmed the existence of the neutron
 - No charge
 - The same mass as a proton

Particle	Symbol	Relative Electrical Charge	Relative Mass Proton = 1	Actual Mass (g)
Electron	e^-	1-	1/1840	9.11×10^{-28}
Proton	p^+	1+	1	1.67×10^{-24}
Neutron	n^0	0	1	1.67×10^{-24}

- The Atomic Nucleus
 - 1911 Ernest Rutherford set about to determine the structure of the atom – sort of hard since you can not see an atom
 - Shot alpha particles at a sheet of gold foil (alpha particle is a helium atom without it's two electrons (2+ charge))
 - Since it was thought that electrons, neutrons and protons were evenly distributed in an atom and the alpha particle would pass through with a small deflection.
 - Some went straight through, some had a small deflection, some had a large deflection and some were bounced back
 - Conclusion – atom is mostly empty space (allowing the alpha particles to pass through. The positive charge and almost all the mass are concentrated in a small region that accounted for the large deflection of some of the particles)
 - This region is know as the nucleus – the core of an atom
 - Composed of protons and neutrons
 - Small compared to the size of an atom

Distinguishing Between Atoms

Objective: How the atomic number identifies an element; Use atomic number and mass number of an element to find the numbers of protons, electrons and neutrons; How isotopes differ, why atomic masses of elements are not whole numbers; Calculate the average atomic mass of an element from isotope data

- Atomic Number
 - All atoms are composed of a nucleus of protons and neutrons surrounded by a large cloud of electrons – how do atoms of different elements differ?

Atoms of the First Ten Elements

Composition of nucleus						
Name	Symbol	Atomic #	Protons	Neutrons*	Mass #	# of electrons

Hydrogen	H	1	1	0	1	1
Helium	He	2	2	2	4	2
Lithium	Li	3	3	4	7	3
Beryllium	Be	4	4	5	9	4
Boron	B	5	5	6	11	5
Carbon	C	6	6	6	12	6
Nitrogen	N	7	7	7	14	7
Oxygen	O	8	8	8	16	8
Fluorine	F	9	9	10	19	9
Neon	Ne	10	10	10	20	10

* Number of neutrons is most abundant isotope

- As can be seen, the number of protons always equals the number of electrons to be electrically neutral.
- Atomic Number = number of protons = number of electrons for a neutral atom
- Mass Number
 - Most of the mass of an atom is in the nucleus (protons & neutrons)
 - Mass Number is the total number of protons & neutrons in an atom
 - Proton & neutron have been assigned a mass number of 1
 - He has 2 protons & 2 neutrons, mass number is 4
 - C has 6 protons & 6 neutrons, mass number is 12
 - The difference between the mass number and the atomic number is the number of neutrons (# of Neutrons = mass number – atomic number)
 - H mass number of 1, atomic number of 1, neutrons = 1-1 = 0
 - O mass number of 16, atomic number of 8, neutrons = 16-8=8
 - Shorthand for showing atomic number and mass number
 - $\begin{matrix} \text{mass_number} \\ \text{atomic_number} \end{matrix} \text{Symbol}$
 - ${}^8_{16}\text{O}$ ${}_{26}^{56}\text{Fe}$ ${}_{92}^{238}\text{U}$
- Isotopes
 - Neutral atoms of a single element will always have the same number of protons and electrons – however the number of neutrons can vary - **isotopes**
 - Example – Neon-20, Neon-21, Neon-22

Neon –20	Neon-21	Neon-22
10 protons	10 protons	10 protons
10 neutrons	11 neutrons	12 neutrons
10 electrons	10 electrons	10 electrons
 - Chemical behavior is determined by electrons & protons

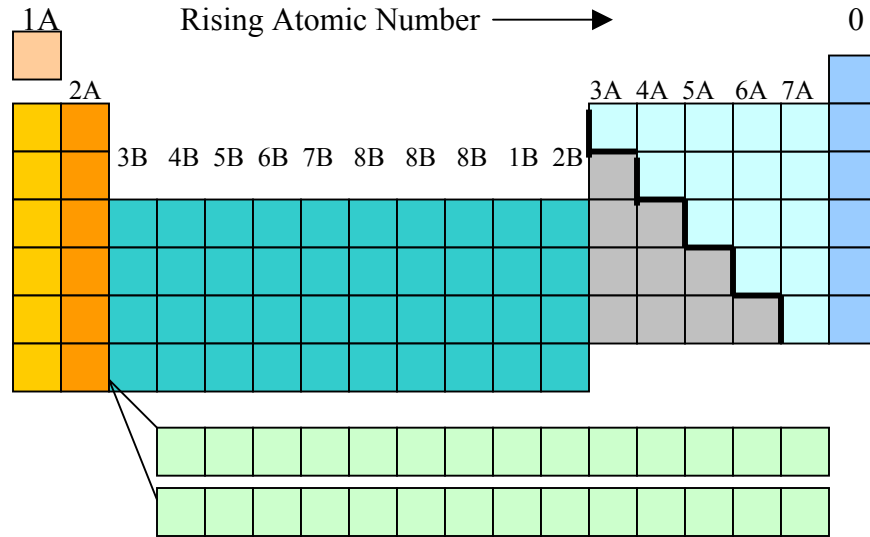
- Isotopes contradict Dalton's theory, since all atoms of an element are not identical
- Atomic Mass
 - Most of the mass of an atom comes from the nucleus – neutrons & protons
 - Scientists have been able to determine the mass of an atom since the 1920's
 - Fluorine – 3.55×10^{-23} g
 - Arsenic – 1.244×10^{-22} g
 - Such numbers are too small to work with
 - Carbon-12 was chosen as the standard and given an atomic mass of 12 (6 neutrons & 6 protons make up the mass – electrons are negligible)
 - Carbon-12 has 12 amu
 - 1 amu = 1/12 Carbon-12 and since there are 6 neutrons and 6 protons in a Carbon-12 atom, 1 proton or 1 neutron has a value of about 1 amu
 - If a proton and neutron have a mass of 1 amu, why do elements not have whole number atomic mass?
 - Isotopes – elements naturally occur with different isotopes
 - Chlorine
 - Chlorine-35 accounts for 75% of all chlorine, Chlorine-37 is the other 25%
 - Taking the weighted average of the two chlorines we find the atomic mass is 35.5 amu
 - Atomic Mass of an element is a weighted average mass of the atoms in a naturally occurring sample of the element
 - Weighted average mass reflects the mass and relative abundance of the isotopes as they naturally occur
- Periodic Table









Objective: Describe the origin of the periodic table; Identify the position of groups, periods, and the transition metals in the periodic table

- In the mid-1800's 70 elements had been discovered
- Russian Dmitri Mendeleev was the first person to arrange the elements in columns so that elements with similar properties were side by side
 - For missing elements, blanks were left and predictions were made of their physical and chemical properties – which turned out to be true
- In 1913 a Brit, Henry Moseley determined the atomic number (# protons) of the elements and arranged them in a table by atomic number instead of atomic mass – still holds today
- Horizontal rows are called **periods**
 - 7 periods
 - 2 elements (period 1) to 32 elements (period 6)

- Properties of a period change as you move across
- Pattern of properties repeat
 - **Periodic Law** – when elements are arranged by increasing atomic number, there is a periodic repetition of chemical and physical properties
 - Elements with similar chemical and physical properties end up in the same column
- Vertical rows are called a **group**
 - Groups have similar chemical and physical properties
 - Groups are identified by a number and either an **A** or **B**
 - Group 1A – all but H react vigorously with water
 - Group 2A, skip across to Group 3A – 7A
 - Group A elements are **representative elements**
 - Wide range of physical and chemical properties
- Group A can be divided into 3 broad classes
 - Metals
 - Highly electrically conductive
 - High luster (shine) when clean
 - Ductile – can be drawn into wire
 - Malleable – beat into thin sheets
 - Most all elements on the far left (except H)
 - Group 1A are **alkali metals**
 - Group 2A are **alkaline earth metals**
 - Most of the Group B elements are metals
 - **Transition metals**
 - Cu, Ag, Au, Fe
 - **Inner transition metals**
 - Two rows below main body of table
 - Rare earth metals
 - 80% of all elements are metals
 - only one metal is not solid at room temperature – Mercury
 - Non-Metals
 - Are in the upper right of the table
 - Poor conductors of electricity
 - Generally are non-lustrous
 - Some are gaseous at room temperature, some are brittle solids
 - Two special groups of non-metals
 - **Halogens** – group 7A
 - **Noble Gases** – group 0
 - Also know as inert gases since they rarely undergo chemical reactions
 - Metalloids

- Elements that border the staircase
- Have properties of both metals & non-metals
 - Silicon, Germanium



	Hydrogen		Other Metals
	Alkali Metals		Non- Metals
	Alkaline Earth Metals		Noble Gases
	Transition Metals		Inner Transition Metals (Rare Erath Metals)

