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Chapter 4 – Atomic Structure

Section 4.1 – Defining the Atom

1. Models of the Atom
   a. Atom
      i. Smallest particle of an element that retains its identity in a chemical reaction
      ii. Comes from the Greek word *atomos,* which means "indivisible"
   b. Democritus believed that atoms were tiny, unchangeable, indivisible, and indestructible.
      i. No experimental support
      ii. Didn't explain chemical behavior
   c. Dalton used experimental methods and transformed Democritus's ideas on atoms into a scientific theory.
      i. Dalton's Atomic Theory
         1. All elements are composed of tiny indivisible particles called atoms.
         2. Atoms of the same element are identical. Atoms of any one element are different from those of any other element.
         3. Atoms of different elements can physically mix together or can chemically combine in simple whole-number ratios to form compounds.
         4. 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.
   d. Radius of most atoms is between 5×10^{-11} m and 2×10^{-10} m
      i. Individual atoms are observable with scanning tunneling microscopes

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Section 4.2 – Structure of the Nuclear Atom

2. While most of Dalton's atomic theory is still accepted, it is now known that atoms are divisible. Atoms can be broken down into subatomic particles: electrons, protons, and neutrons.
   a. Electrons
      i. Negatively charged subatomic particles
      ii. Discovered by Thomson by passing electric current through gases at low pressure.
         1. Cathode ray tube – electrons travel as a ray from the cathode (negatively charged) towards the anode (positively charged).
         2. Cathode ray was deflected by using a magnet and by electrically charged plates.
            a. Thomson hypothesized that the cathode ray was a stream of tiny negatively charged particles moving at high speed.
      iii. Millikan carried out experiments to find the quantity of charge carried by an electron. He used Thomson's charge to mass ratio of an electron to calculate the mass of the electron.
      iv. Properties:
         1. Symbol e^-
         2. Charge -1
         3. Actual mass 9.11×10^{-28}
         (1840 times smaller than a proton)
   b. Protons
      i. Positively charged subatomic particles
      ii. Goldstein observed a cathode-ray tube and found rays traveling in the opposite direction. He called them Canal rays and concluded that they were made of positive particles.

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iii. Properties:

1. Symbol p^+
2. Charge +1
3. Actual mass 1.67×10^{-24}

c. Neutrons

i. Subatomic particles with no charge but with a mass nearly equal to a proton

ii. Properties

1. Symbol n^0
2. Charge 0
3. Actual mass 1.67×10^{-24}

iii. Chadwick confirmed the existence of the neutron

3. Atomic Nucleus

a. Thomson's Plum Pudding Model

i. Believed that electrons were evenly distributed throughout an atom filled uniformly with positively charged material

ii. Electrons were stuck into a lump of positive charge, like raisins stuck in dough

b. Rutherford's Gold-Foil Experiment

i. Aimed a beam of alpha particles at a sheet of gold foil surrounded by a fluorescent screen. Most of the particles passed through the foil with no deflection at all. A few particles were greatly deflected. Rutherford concluded that most of the alpha particles pass through the gold foil because the atom is mostly empty space.

ii. The mass and positive charge are concentrated in a small region of the atom. He called this area the nucleus. Particles that approach the nucleus closely are greatly deflected.

c. Rutherford's Atomic Model

i. Atom is mostly empty space.

ii. Nucleus – tiny central core of an atom. Composed of protons and neutrons.

iii. Electrons are distributed around the nucleus and occupy almost all the volume of the atom.

Section 4.3 – Distinguishing Among Atoms

4. Preview of the Periodic Table
   a. The Periodic Table allows you to easily Compare the properties of one element (or a group of elements) to another element (or group of elements).
   b. Elements are organized intro groups based on similar chemical properties
      i. Periods
         1. Horizontal rows (go across) ⟶
         2. Labeled from 1 to 7
      ii. Groups
         1. Also called families
         2. Vertical columns (go up + down) ↓
         3. Labeled from 1 to 18
   c. Groups on the Periodic Table
      i. Group 1 Alkali metals
      ii. Group 2 Alkaline earth metals
      iii. Groups 3-12 Transition metals
      iv. Group 13 Boron Family
      v. Group 14 Carbon Family
      vi. Group 15 Nitrogen Family
      vii. Group 16 Oxygen Family
      viii. Group 17 Halogens
      ix. Group 18 Noble Gases
5. Atomic Number
   a. Elements are different because they contain different number of protons.
   b. The atomic number of an element is the number of protons in the nucleus of an atom of that element.
   c. Atoms have a neutral charge. The number of protons equals the number of electrons.

d. "Top" number on the Periodic Table.

e. The number increases across the Periodic Table, as each element has more protons than the one to its left.

6. Mass Number

a. Most of the mass of an atom is concentrated in the nucleus.

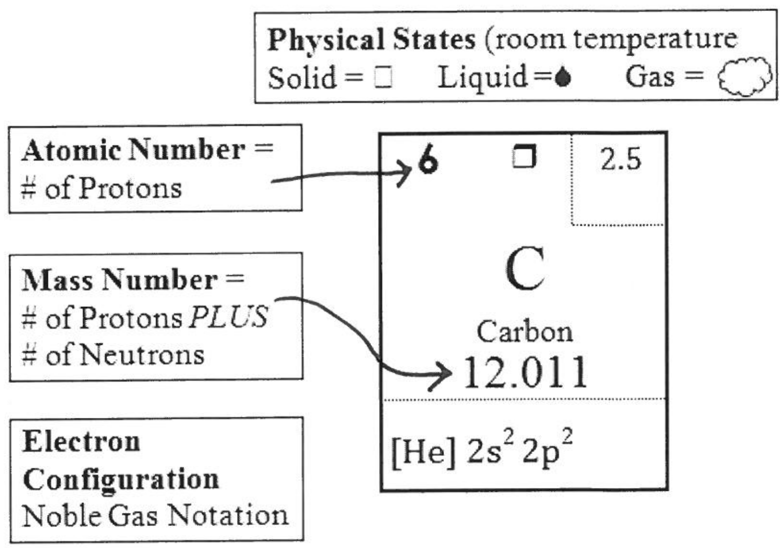
b. The mass number is the total number of protons and neutrons.

c. "Bottom" number on the Periodic Table

Mass # = protons + neutrons

Protons = Atomic #

Neutrons = Mass # - protons



7. Isotopes

a. Isotopes are atoms that have the same number of protons but different numbers of neutrons (different mass)

b. Isotopes are chemically alike because they have identical numbers of protons and electrons, which are the subatomic particles responsible for chemical behavior.

c. 3 isotopes of Hydrogen

i. Hydrogen, hydrogen-1
1. 1 proton, 0 neutrons, 1 electron

ii. Deuterium, hydrogen-2
1. 1 proton, 1 neutron, 1 electron

iii. Tritium, hydrogen-3
1. 1 proton, 2 neutrons, 1 electron

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8. Ions
   a. Ions are atoms with a charge
      i. Examples: Al^{+3} F^{-1} Ba^{+2} O^{-2}
   b. Caused by gaining or losing electrons
      i. Gaining electrons
         1. Gives an atom a negative charge, as it is getting more negative things.
      ii. Losing electrons
         1. Gives an atom a positive charge, as it is getting rid of negative things
   c. Positive ions are called cations
   d. Negative ions are called anions
   e. Example:
      i. Ca^{+2}
         1. Protons = atomic # 20
         2. Electrons = protons – charge 18 20 −(+2)
         3. Neutrons = mass – protons 20 40 - 20
      ii. P^{-3}
         1. Protons = atomic # 15
         2. Electrons = protons – charge 18 15 −(-3)
         3. Neutrons = mass – protons 16 31 −15

9. The one constant is the proton. It will never change. It always equals the atomic number.
   a. Isotope – change in neutrons. (Same atom but with different masses)
   b. Ions – change in electrons. (Same atom but with either a +/− charge)

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10. Notation

a. Complete chemical symbol

i. Shorthand notation using the chemical symbol, atomic number and mass number

ii. Mass number is written as a superscript

iii. Atomic number is written as a subscript

1. This format allows for easy calculation of the number of neutrons

$${}^{\text{mass \#}}_{\text{atomic \#}}X^{\text{charge}}$$

iv. Examples:

1. Silver

a. Protons 47
b. Neutrons 61
c. Electrons 47

Mass = 47 + 61

${}^{108}_{47}Ag$

2. Tungsten

a. Protons 74
b. Neutrons 110
c. Electrons 74

Mass = 74 + 110

${}^{184}_{74}W$

3. Nitrogen

a. Protons 7
b. Neutrons 7
c. Electrons 7

Mass = 7 + 7

${}^{14}_{7}N$

b. Hyphen notation

i. Write the full name of the element, then a hyphen, then the mass number (number of protons and neutrons)

ii. Examples:

1. Silver-108
2. Tungsten-184
3. Nitrogen-14

11. Atomic Mass Unit

a. Defined as one twelfth of the mass of a carbon-12 atom

b. The atomic mass of an element is a weighted average mass of the atoms in a naturally occurring sample of the element. A weighted average mass reflects both the mass and the relative abundance of the isotopes as they occur in nature.

i. To calculate the atomic mass of an element, multiply the mass of each isotope by its natural abundance (expressed as a decimal) and then add the products.

ii. Chlorine has two isotopes: chlorine-35 and chlorine-37

1. Chlorine-35 has an atomic mass of 34.969 and is found in nature 75.77%

2. Chlorine-37 has an atomic mass of 36.966 and is found in nature 24.23%

i. (34.969)(0.7577) = 26.4960113

ii. (36.966)(0.2423) = + 8.9568618

35.4528731

This is the average atomic mass

12. Calculating Subatomic Particles:

a. Protons = Atomic number

b. Electrons = # of protons – (charge)

c. Neutrons = mass number – atomic number

d. Mass # = protons + neutrons

e. Charge = protons – electrons

Example Chlorine Atomic # 17
Mass # 35
Charge of -1

$^{35}_{17}Cl^{-1}$ p = 17 n = 18 e = 18