CHAPTER 2

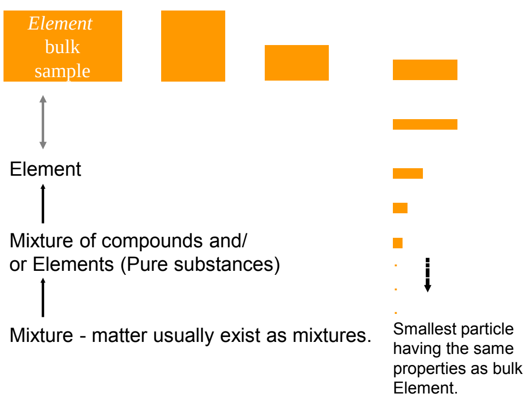
Atoms, Ions, and Compounds

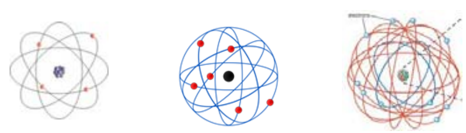
Chemists' approach to the understanding of matter:

should explain the properties of macroscopic quantities of matter from the nano-scale point of view.

'nanoscale' - the molecular/atomic *size*

atoms
molecules (combinations of atoms)
ions (electrically charged 'atomic/molecular' species.





Atom

Smallest particle of an element having the same chemical properties as bulk quantities of that element.

Atomic Structure:

What is an atom made of?
How does an atom look like (interior)?

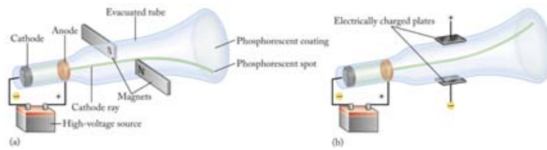
A *series of experiments and observations* led to the current 'model' of the atom.

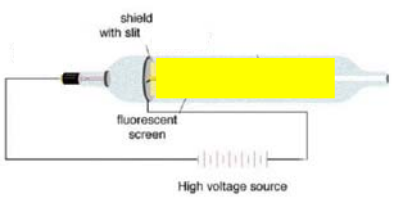
To interpret the experiments, recall that:

Like electrical charges repel and unlike electrical charges attract.

Atomic Structure: Electrons

- J. J. Thomson (1897):
  - Beam from cathode ray tube deflected toward positively charged plate. Cathode ray negatively charged
  - Atoms contain negatively charged particles with a constant mass-to-charge ratio.





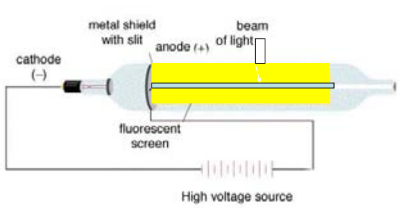
Cathode Ray Experiment:

Evacuated tube with metal plates at ends (electrodes; anode and cathode).

Anode has a slit cut off at its center.

A fluorescent screen placed at the center.

A high voltage applied across the electrodes. Negative and positive potentials on electrodes.



Cathode Ray Experiment:

Observations:
A bright ray appeared on the fluorescent screen.

Change the size of slit; *size of beam on screen changes accordingly.*

Reverse the Polarity
Observation: Ray disappears.

Cathode Ray Experiment:

Interpretation
Whatever was formed within the electrodes *leaks out* of the slit and creates a light beam when struck on screen.

Whatever formed within electrodes move from (-) to (+) (as a beam). The 'beam' originates from the negative electrode (cathode)

Cathode "ray".

Cathode Ray Experiment:

Whatever formed must be negatively charged.

Change the metal (material) of the cathode. *Same results.*

*All materials contain these negatively charged species - **ELECTRONS**.*

Electrons are a basic component of matter, (atoms).

Cathode 'rays' = beam of moving, negatively charged particles.

Charged particles interact with *both* magnetic and electric fields.

Consider the electric field case.

By studying the degree of bend when the field is turned ON, the

charge/mass ratio of an electron calculated.

$(q/m)_e = 1.76\times10^8$ C/g

Movie1 Movie2 Movie3

Radioactivity and the Nuclear Atom

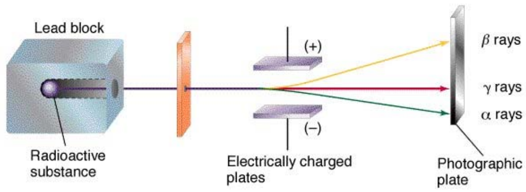
- Henri Becquerel (1896):
  - Some materials produce invisible radiation, consisting of charged particles.
- Radioactivity:
  - Spontaneous emission of high-energy radiation.
    - beta particles (β, high energy electrons)
    - alpha particles (α, +2 charge, mass = He nucleus)

Radioactivity:

Some heavy elements spontaneously disintegrate into lighter elements.

In the process of the decay, they emit smaller decay particles as a 'radiation' beam.

Analysis of the radiation beam revealed that it contains three parts.



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The decay particles gives us a glimpse as to what the original heavy atom (and therefore atoms) are made of.

Premise: Decay particles were *in* the atom to start with.

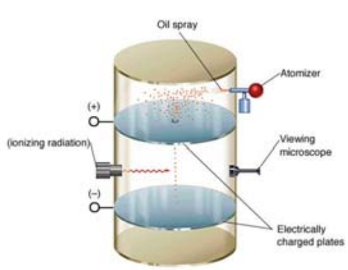
β particles = cathode rays (electrons)
electrical charge negative (charge -1)
(a component of atoms - already established)

α charge opposite to that of electrons (positive) and twice in magnitude (charge +2).
much heavier than electrons

γ massless, (= X-rays)
high energy ‘light’
(electro-magnetic radiation)

These rays penetrate through matter

γ > β > α



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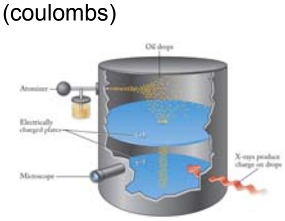
Mass of an Electron

Robert Millikan (1909):

- Determined the mass and charge of an electron with his oil-droplet experiment.
- $e^- = -1.602 \times 10^{-19}$ C (coulombs)
- $m_e = 9.109 \times 10^{-28}$ g



By studying the motion of oil droplets in the electrical field

charge of an electron calculated.

$(q)_e = -1.60\times10^{-19}$ C (= -1, for convenience)

Then, substituting in the $(m/q)_e=5.60\times10^{-9}$ g/C

$m_e=9.11\times10^{-28}$g !!

$2\ \beta + \alpha \rightarrow$ helium atom (Rutherford)

So atoms contain positively and negatively charged particles.

A negatively charged particle = electron
A positively charged particle = proton

Helium atom has two electrons and two protons; overall charge zero, electrically neutral.

In any atom; #electrons = # protons

Lightest element = Hydrogen

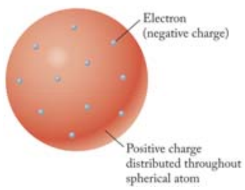
Mass of one H atom $\approx 1800 \times m_e \sim 1.63 \times 10^{-24}$g

Thompson Model of Atom:

Uniform, positively charged sphere with electrons embedded in it.

Thomson Model of the Atom

- Plum-Pudding Model:
  - Matter is electrically neutral. Total # positive particles = total # electrons. e^- distributed throughout diffuse, positively charged sphere.



Subsequent work showed that atoms contain electrically neutral heavy particles as well - neutrons.

He = 2 electrons + (2 protons + 2 neutrons)
↓
α particle

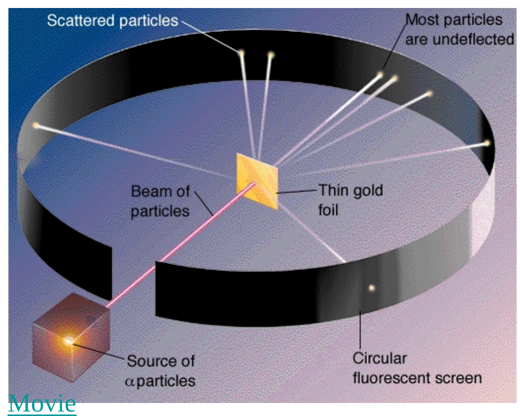
Particle	Charge	Mass (amu)
Proton	Positive +1	1.0073 ~1
Neutron	None	1.0087 ~1
Electron	Negative -1	5.486×10^{-4}

1 amu = 1.66054×10^{-24}g
charge of 1 = 1.60217×10^{-19} C

Rutherford's alpha particle scattering experiment

A beam of alpha particles aimed at a thin metal (Au) foil.

Fate of the particle beam studied.



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Observations:

1. Most particles un-deflected.

2. A few particles deflected.

3. Even lesser number of particles reflected back.

Interpretation

- Most alpha particles do not encounter a resistance to their movement thro' the foil.
  - Most of the atomic volume is soft.
- Very small number of particles are reflected
  - Hard part of the atom is small.
- Deflections - due to particles pushed away
  - +ve alpha particles pushed aside by the +ve part of the the atom and that part is small in volume.

- Of the three types of particles, lighter - electron
  - softer part of the atom is made of light electrons, and they occupy most of the volume and is -vely charged.
- Hard part of the atom is made of heavy particles
  - protons and neutrons and they occupy a very small volume.
  - overall the hard part of the atom is +vely charged.

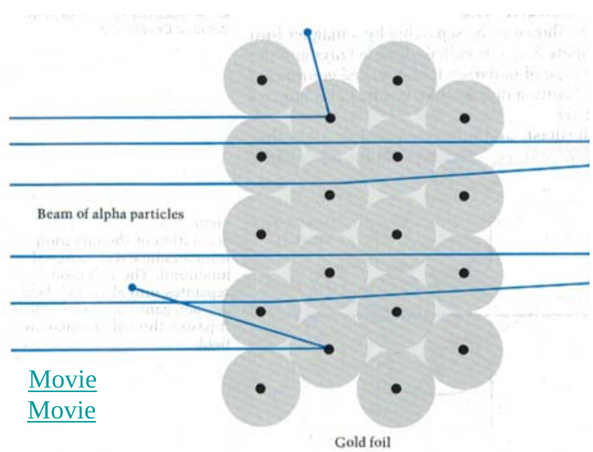
<u>Rutherford's Nuclear Atom</u>:

Atom: heavy positively charged small region, soft negatively charged large region.
#protons = # electrons; atom overall neutral.

Nucleus (hard): protons + neutrons
virtually all mass is in the nucleus.

Rest of the atom is a soft sphere.

Radius of the nucleus = (1/100,000) atom radius.



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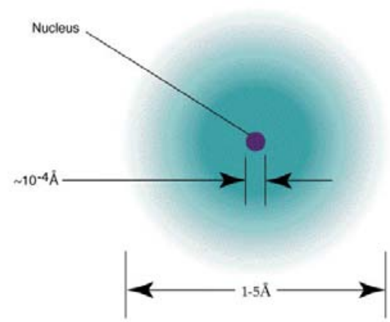
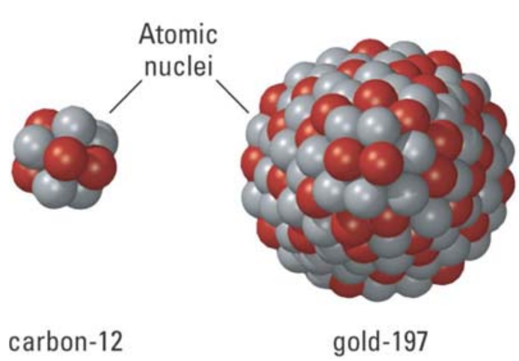
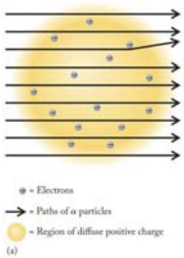




<u>Atomic Structure: The Nucleus</u>

- Rutherford's Experiment:
  - Bombarded a thin gold foil with *α* particles to test Thomson's model of the atom.
  - Theory predicted that the *α* particles would travel through the foil with little or no deflection.
  - Results indicated presence of dense particle within the atom.

Rutherford's Experiment

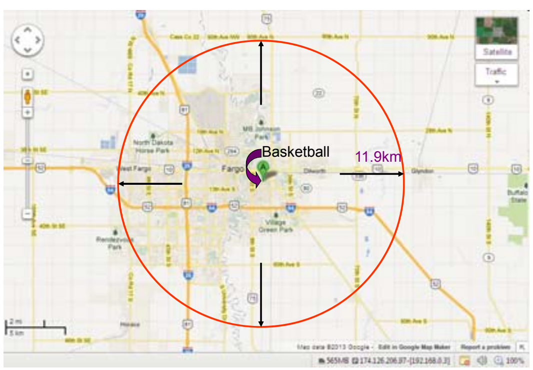


a) Expected results from "plum-pudding" model.

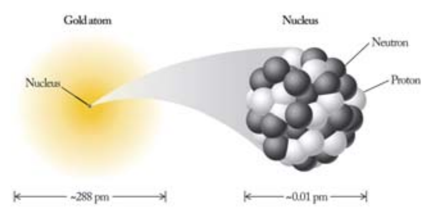
b) Actual results.

The Nuclear Atom

- The nucleus:
  - Contains all the positive charge and nearly all the mass in an atom.
  - Is about 1/10,000 the size of the atom.
  - Consists of two types of particles:
    - » proton: positively charged particle
    - » neutron: neutral particle



Nucleus: contains both positively charged particles (protons) and neutral particles (neutrons).



Atomic Mass Units

- Atomic Mass Units (amu)
  - A relative scale to express the masses of atoms and subatomic particles.
  - Scale is based on the mass of 1 atom of carbon:
    - » 6 protons + 6 neutrons = 12 amu.
  - 1 amu = 1 Dalton (Da)

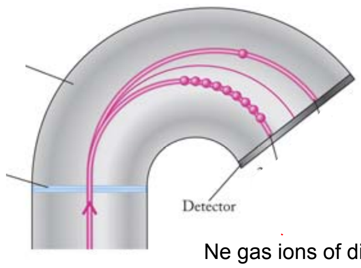
The Nuclear Atom: Summary

TABLE 2.1 Properties of Subatomic Particles

		MASS		CHARGE	
Particle	Symbol	In Atomic Mass Units (amu)	In Grams (g)	Relative Value	Charge (C[a])
Neutron	${}^{1}_{0}n$	1.00867 ≈ 1	1.67493×10^{-24}	0	0
Proton	${}^{1}_{1}p$	1.00728 ≈ 1	1.67262×10^{-24}	1+	$+1.602 \times 10^{-19}$
Electron	${}^{0}_{-1}e$	$5.485799 \times 10^{-4} \approx 0$	9.10939×10^{-28}	1−	-1.602×10^{-19}

[a] The *coulomb* (C) is the SI unit of electric charge. When a current of 1 ampere (see Table 1.2) passes through a conductor for 1 second, the quantity of electric charge that moves past any point in the conductor is 1 C.

Aston's Positive-Ray Analyzer



Ne gas ions of different masses strike detector in different locations.

- Positive Ray Analyzer Results:
  - Two different kinds of neon gas atoms existed:
    - » 90% = 20 amu
    - » 10% = 22 amu
  - Aston proposed theory of "isotopes".
- Isotopes:
  - Atoms of the same element (same number of protons) but different numbers of neutrons (different mass).

Symbol of Isotopes

Atomic Mass (A) = total number of "nucleons" (protons, neutrons) in the nucleus.

$^{A}_{Z}X$

Elemental Symbol = a one- or two-letter symbol to identify the type of atom.

Atomic Number (Z) = the number of protons in the nucleus; determines the identity of the element.

Practice: Isotopic Symbols

Use the format ^{A}X to write the symbol for the nuclides having 26 protons and 30 neutrons.

Average Atomic Masses

- Natural Abundance:
  - Relative proportion of a given isotope compared to all the isotopes for the element found in a natural sample.
  - Expressed as percent.
- Average Atomic Mass:
  - Weighted average mass of natural sample of an element, calculated by multiplying the natural abundance of each isotope by its exact mass in amu and then summing these products.

Weighted Average Example

Neon is found in three isotopes in nature.

Isotope	Mass (amu)	Natural Abundance (%)
Neon-20	19.9924	90.4838
Neon-21	20.9940	0.2696
Neon-22	21.9914	9.2465

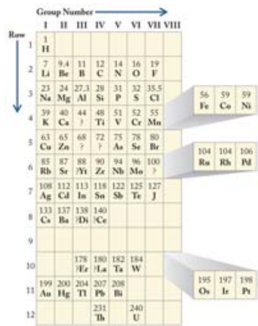
Average atomic mass of neon:

$(19.9924 \times 0.904838) + (20.99395 \times 0.002696)$
$+ (21.9914 \text{ x } 0.092465) = 20.1797$ amu

Mendeleev's Periodic Table

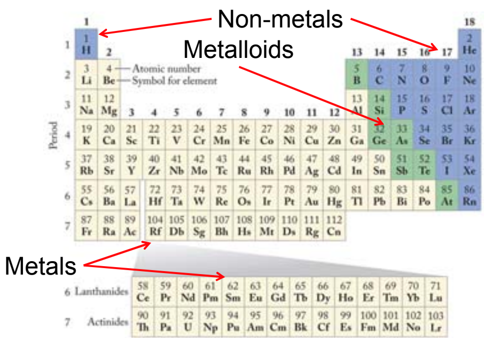
Dmitrii Mendeleev (1872):

- Ordered elements by *atomic mass*.
- Arranged elements in columns based on similar chemical and physical properties.
- Left open spaces in the table for elements not yet discovered.

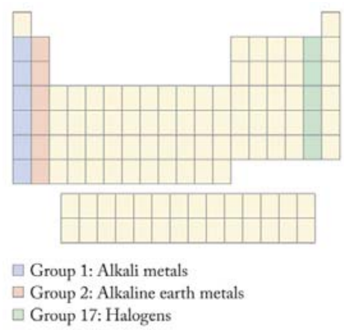


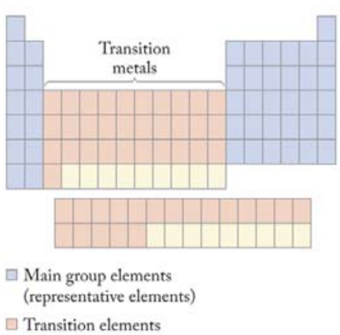
The Modern Periodic Table

- Also based on a classification of elements in terms of their physical and chemical properties.
- Horizontal rows: called **periods** (1 → 7).
- Columns: contain elements of the same family or **group** (1 →18).
- Several groups have names as well as numbers.
- Ordered elements by *Atomic Number, Z*.



Groups of Elements





Broad Categories of Elements

- Metals (left side and bottom of the table)
  - Shiny solids; conduct heat and electricity; are malleable and ductile.
- Nonmetals (right side and top of the table)
  - Solids, liquids, and gases; non-conductors; solids are brittle.
- Metalloids (between metals/nonmetals)
  - Shiny solids (like metals); brittle (like nonmetals); semiconductors.

The Composition of Compounds

- Law of Multiple Proportions
  - If two elements can combine to form *more than one compound*, the mass of Y that will react with a given mass of X to form the compounds can be expressed as a ratio of small whole numbers.
- Examples: NO, NO_2, N_2O, N_2O_5, etc.

Molecular Compounds

- Molecular Compounds
  - Composed of atoms held together by *covalent bonds*.
- Covalent Bonds
  - *Shared pairs of electrons* that chemically bond atoms together.
- Molecular Compounds Composed of *Nonmetals*.

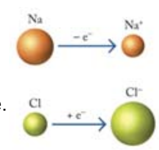
- Molecular Formulas:
  - Shows the exact number and elements present in one molecule of a compound (*e.g.*, H_2O, CO_2).
- Empirical Formula:
  - Gives the *simplest/lowest whole-number* ratio of elements in a compound.
  - Example: Glucose
    - » molecular formula = $C_6H_{12}O_6$
    - » empirical formula = CH_2O

Ionic Compounds

- Ionic compounds:
  - *Charged particles* (ions) formed by transfer of electrons between atoms.
  - Ions held together by *electrostatic forces*.

Cations = Ions with positive charge.

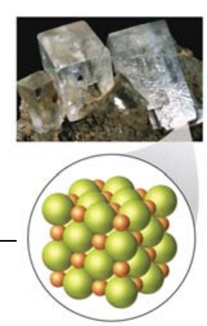
Anions = Ions with negative charge.



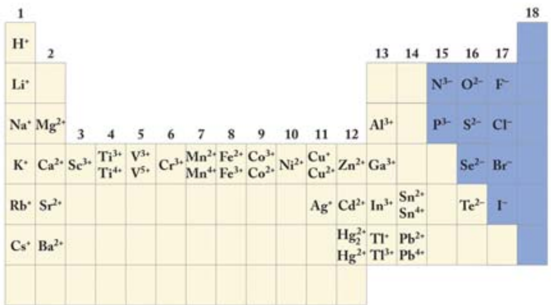
- Ionic compounds are made of a *metal and a nonmetal.*
  - Metals form cations; nonmetals form anions.
  - Charges on ions depend on location in the periodic table.
    - » *e.g.*, Group 1 Metals = +1; Halogens = −1

- Formula Unit - Ionic Compounds
  - Smallest electrically neutral unit within the crystal of the compound.

e.g., NaCl



Charges on Ions



Naming Compounds

- Binary Molecular Compounds (*e.g.*, SO_3):
  - Compounds consisting of two nonmetals:
    - » First element in the formula is named first.
      - S = sulfur
    - » Second element name is changed by adding suffix –ide.
      - O = oxygen → oxide
    - » Use prefixes to identify quantity of atoms (see Table 2.2).
      - SO_3 = sulfur trioxide

Rules for Using Prefixes

1. Do not use the prefix mono- when naming first element:

   SO_3 ~~mono~~sulfur trioxide

2. Prefixes ending with *o* or *a* are modified when used with elements beginning with vowels:

   P_4O_{10} tetraphosphorus dec~~a~~oxide

TABLE 2.2 Naming Prefixes for Molecular Compounds

one	*mono-*
two	*di-*
three	*tri-*
four	*tetra-*
five	*penta-*
six	*hexa-*
seven	*hepta-*
eight	*octa-*
nine	*nona-*
ten	*deca-*

Practice: Naming Binary Molecular Compounds

1. Name the following compounds:
   a) CCl_4
   b) P_2N_5

2. Give the correct chemical formula for the following compound names:
   a) Sulfur trioxide
   b) Tetraphosphorus decoxide

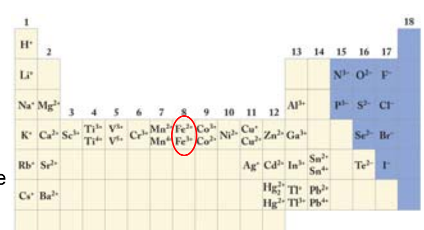
Binary Ionic Compounds

- Binary ionic compounds consist of cations (usually metals) and anions (usually nonmetals). (*e.g.*, $MgCl_2$)
  - The cation is named first using the elemental name.
    - » Mg = magnesium
  - The anion is named by adding the –ide suffix to the name of the element.
    - » Cl = chlorine → chloride
  - Formulas for ionic compounds must always be neutral: Mg^{2+} + (Cl^{-}) × 2

For metals that form cations with different charges, a Roman numeral is added to indicate the charge of the cation.

$FeCl_2$: Iron(II) chloride

$FeCl_3$: Iron(III) chloride



Practice: Ionic Compunds

1. Write the name of the following compounds:
   a) NaCl
   b) $CrCl_3$
2. Write the chemical formula of the following compounds:
   a) Zinc nitride
   b) Copper(I) oxide

Common Polyatomic Ions

- Acetate $C_2H_3O_2^-$
- Carbonate CO_3^{2-}
- Perchlorate ClO_4^-
- Nitrate NO_3^-
- Sulfate SO_4^{2-}
- Chromate CrO_4^{2-}

(Memorize Table 2.3 on Page 62)

TABLE 2.3 Names and Charges of Some Common Ions

Name	Chemical Formula	Name	Chemical Formula
Acetate	CH_3COO^-	Hydrogen phosphate	HPO_4^{2-}
Ammonium	NH_4^+	Hydrogen sulfite or bisulfite	HSO_3^-
Azide	N_3^-	Hydroxide	OH^-
Bromide	Br^-	Nitrate	NO_3^-
Carbonate	CO_3^{2-}	Nitride	N^{3-}
Chlorate	ClO_3^-	Nitrite	NO_2^-
Chloride	Cl^-	Oxide	O^{2-}
Chromate	CrO_4^{2-}	Perchlorate	ClO_4^-
Cyanide	CN^-	Permanganate	MnO_4^-
Dichromate	$Cr_2O_7^{2-}$	Peroxide	O_2^{2-}
Dihydrogen phosphate	$H_2PO_4^-$	Phosphate	PO_4^{3-}
Disulfide	S_2^{2-}	Sulfate	SO_4^{2-}
Fluoride	F^-	Sulfide	S^{2-}
Hydride	H^-	Sulfite	SO_3^{2-}
Hydrogen carbonate or bicarbonate	HCO_3^-	Thiocyanate	SCN^-

Practice: Polyatomic Ions

1. Write the names of the following compounds:
   a) $Cr(ClO_4)_3$
   b) NH_4NO_3
2. Write the chemical formulas for the following compounds:
   a) Lithium bicarbonate
   b) Calcium hypobromite

Naming Binary Acids

- Binary acids:
  - Contain hydrogen and a monoatomic anion (*e.g.*, Cl^-, S^{2-}).
  - Most common binary acids are halogen (*e.g.*, HCl, HBr).
  - Acid names:
    - The prefix “hydro” + the halogen base name + the suffix “ic” + the word acid
  - Example: HBr—hydrobromic acid.

Oxy Anions & Related Acids

If oxoanion name ends in:	corresponding acid ends in:
–ate	–ic
–ite	–ous

TABLE 2.4 Oxoanions of Chlorine and Their Corresponding Acids

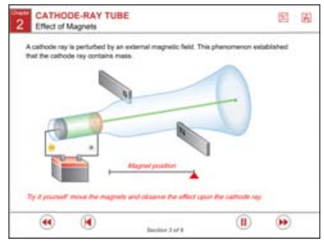
Ions		Acids	
ClO^-	hypochlorite	HClO	hypochlorous acid
ClO_2^-	chlorite	$HClO_2$	chlorous acid
ClO_3^-	chlorate	$HClO_3$	chloric acid
ClO_4^-	perchlorate	$HClO_4$	perchloric acid

Practice: Naming Compounds and Acids

Identify each of the following as a molecular compound, an ionic compound, or an acid. Name or give formulas for the compounds.

- $K_2Cr_2O_7$
- Na_3N
- NO_2
- H_2CrO_4
- Sodium carbonate
- Sulfurous acid
- Iron(II) phosphate

ChemTour: Cathode Ray Tube



launch this ChemTour

This ChemTour explores the effects of magnetic and electric fields and cathode rays.

ChemTour: Millikan Oil Drop Experiment

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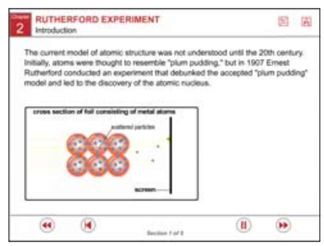
This ChemTour recreates the experimental procedure used by Millikan to determine the charge of an electron.

ChemTour: Rutherford Experiment



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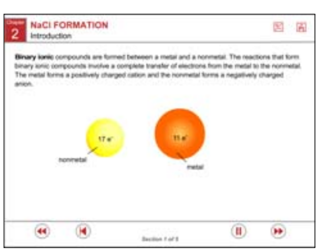
This recreates Rutherford's gold foil experiment, which led to the discovery of the atomic nucleus.

ChemTour: NaCl Reaction



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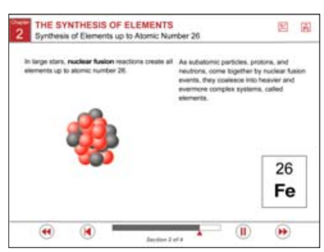
This ChemTour illustrates the process by which a metal and a nonmetal combine to form a binary ionic compound, as seen in the reaction of sodium metal and chlorine gas.

ChemTour: Synthesis of Elements



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This ChemTour animates the neutron capture process and explains how elements are synthesized in stars.

Sample Exercise 2.1

Write symbols in the form $^{A}_{Z}X$ for the nuclides that have (c) 92 protons and 143 neutrons.

- We know the number of protons and neutrons in the nuclei of three nuclides and are to write symbols of the form where *Z* is the atomic number, *A* is the mass number, and X is the symbol of the element.

$$^{A}_{Z}\mathrm{X}$$

Sample Exercise 2.2

The precious metal platinum (*Z* = 78) has six isotopes with these natural abundances:

Symbol	Mass (amu)	Natural Abundance (%)
^{190}Pt	189.96	0.014
^{192}Pt	191.96	0.782
^{194}Pt	193.96	32.967
^{195}Pt	194.97	33.832
^{196}Pt	195.97	25.242
^{198}Pt	197.97	7.163

Use these data to calculate the average atomic mass of platinum.

- We know the masses and natural abundances of each of the six isotopes of platinum and are asked to calculate the average atomic mass.

Sample Exercise 2.2 (cont.)

- we multiply the mass of each isotope by its natural abundance expressed as a decimal, and then add the products together.

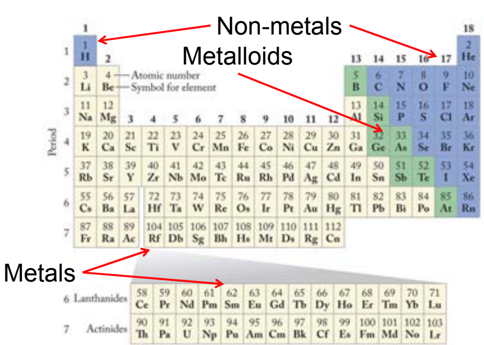
$$\begin{aligned}\text{Average atomic mass} = {} & (189.96\ \text{amu})(0.00014) \\ & + (191.96\ \text{amu})(0.00782) \\ & + (193.96\ \text{amu})(0.32967) \\ & + (194.97\ \text{amu})(0.33832) \\ & + (195.97\ \text{amu})(0.25242) \\ & + (197.97\ \text{amu})(0.07163) \\ & \overline{\quad 195.08\ \text{amu}}\end{aligned}$$

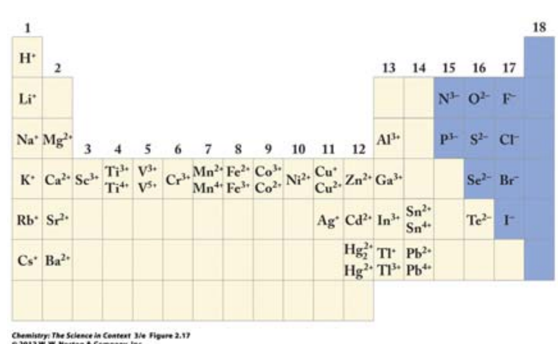
Sample Exercise 2.3

Give the symbol and name of each element:

The metal (X) in the second row that forms a compound with the chemical formula XBr_2

- We are to identify elements based on the locations of their symbols in the periodic table. We are given the row number, a chemical property.





Sample Exercise 2.3 (cont.)

- The group 2 elements form 1:2 compounds with Br and the other halogens, so the cell address is group 2, row 2.
- Be,beryllium

Sample Exercise 2.3 (cont.)

- Each element has a unique location in the periodic table determined by its atomic number, which defines the row it is in, and its reactivity with other elements, which defines the group it is in. We assumed that beryllium was the only metal in the second row that could form a 1:2 compound with bromine. This is a valid assumption because the only other metal in the second row is Li, which is a group 1 element whose compound with Br has the formula LiBr.

Sample Exercise 2.4

Carbon combines with oxygen to form either CO or CO_2 depending on reaction conditions. If 26.6 g of oxygen reacts with 10.0 g of carbon to make CO_2, how many grams of oxygen reacts with 10.0 g of carbon to make CO?

- The two compounds contain the same two elements but in different proportions, so Dalton's law of multiple proportions applies. We have formulas for both compounds, CO and CO_2, and are told that both reactions involve 10.0 g of carbon.

Sample Exercise 2.4 (cont.)

- The ratio of the O atoms to C atoms in CO is 1:1. The ratio of O atoms to C atoms in CO_2 is 2:1.Therefore, half as much oxygen will react with10.0 g of carbon to make CO as reacts with 10.0 g of carbon to make CO_2.
- (26.6 g of oxygen) × ½ 13.3 g of oxygen

Sample Exercise 2.4 (cont.)

- We used these chemical formulas in this exercise to calculate the different masses of oxygen required to react completely with a given mass of carbon to form the two compounds. In actual practice, the reverse is done: chemists analyze the masses of the elements in a compound and use that information to determine its molecular formula.

Sample Exercise 2.5

Identify each of the following compounds as ionic or molecular: (a) sodium bromide (NaBr); (b) carbon dioxide (CO_2); (c) lithium iodide (LiI); (d) magnesium fluoride (MgF_2); (e) calcium chloride ($CaCl_2$).

- We are to distinguish between ionic and molecular compounds based on their names and chemical formulas. In this section we learned that compounds formed by reacting metals with nonmetals tend to be ionic; those that contain only nonmetallic elements are molecular. We can use the periodic table to determine which of the elements in the compounds are metallic and which are nonmetallic.

Sample Exercise 2.5 (cont.)

- NaBr, LiI, MgF_2, and $CaCl_2$ all contain a group 1 or group 2 metal and a group 17 nonmetal. Only CO_2 is composed of two nonmetals.
- (a) NaBr, (c) LiI, (d) MgF_2, and (e) $CaCl_2$ are ionic; (b) CO_2 is molecular.

Sample Exercise 2.5 (cont.)

- In later chapters we will discover that the world of compounds is not so black and white as painted in this exercise. Some covalent bonds have a degree of ionic “character,” and we will explore a way based on the elements’ positions in the periodic table to determine how much ionic character covalent bonds have.

Sample Exercise 2.6

What are the names of the compounds with these chemical formulas: (a) N_2O; (b) N_2O_4; (c) N_2O_5?

- All three compounds are binary nonmetal oxides and hence molecular compounds. Therefore, we use prefixes in the names to indicate the number of atoms of each element present in one molecule.

Sample Exercise 2.6 (cont.)

a. dinitrogen monoxide
b. dinitrogen tetroxide
c. dinitrogen pentoxide

- To avoid back-to-back vowels in the middle of the second terms in all three names, we deleted the last letter of the three prefixes before *oxide*.

Sample Exercise 2.7

Write the chemical formula of (a) potassium bromide, (b) calcium oxide, (c) sodium sulfide, (d) magnesium chloride, and (e) aluminum oxide.

- The name of each compound consists of the name of one main group metal and one main group nonmetal, which tells us that these are binary ionic compounds. To write formulas of ionic compounds, we assign the charges on the ions based on the group numbers of the parent elements.

Sample Exercise 2.7 (cont.)

- Locate each element in the periodic table and predict the charge of its most common ion based on location and group number: K^+, Br^-, Ca^{2+}, O^{2}, Na^-, S^{2-}, Mg^{2+}, Cl^-, and Al^{3+}. If you have difficulty predicting ionic charge, refer to the book. Writing chemical formulas of the compounds is an exercise in balancing positive and negative charges.

Sample Exercise 2.7 (cont.)

- We must balance the positive and negative charges in each compound:

  a. In potassium bromide, the ionic charges are 1+ and 1- (K^+ and Br). A 1:1 ratio of the ions is required for electrical neutrality, making the formula KBr.

  b. CaO.

  c. Na_2S.

  d. $MgCl_2$.

  e. Al_2O_3.

Sample Exercise 2.7 (cont.)

- Different approaches may be used to work out the formulas of ionic compounds. The basic principle is that the sum of the total positive and negative charges must balance to give a net charge of zero. If you had difficulty writing the formula of aluminum oxide, try this shortcut: use the charge on each ion as the subscript for the other ion. Thus the 3+ charge on Al^{3+} becomes a subscript $_3$ after O, and the 2- charge on the oxide ion becomes a subscript $_2$ after Al. The result is Al_2O_3:

$$Al^{3+}O^{2-} \rightarrow Al_2O_3$$

Sample Exercise 2.8

(a) Write the chemical formulas of iron(II) sulfide and iron(III) oxide. (b) Write alternative names for these compounds that do not use Roman numerals to indicate the charge on the iron ions.

- We are to write chemical formulas for two ionic compounds. Because iron is a transition metal, Roman numerals are used to indicate the charges on the iron ions.

Sample Exercise 2.8 (cont.)

- The Roman numerals (II) and (III) indicate that the charges on the iron cations are 2+ and 3+, respectively. Oxygen and sulfur are both in group 16. Therefore the charge on both the sulfide ion and oxide ion is 2-. In the alternate naming system, Fe^{2+} is the *ferrous* ion and Fe^{3+} is the *ferric* ion.

  a. A charge balance in iron(II) sulfide is achieved with equal numbers of Fe^{2+} and S^{2-} ions, so the chemical formula is FeS. To balance the different charges on the Fe3+ and O^{2-} ions in iron(III) oxide, we need three O^{2-} ions for every two Fe^{3+} ions. Thus the formula of iron(III) oxide is Fe_2O_3.

  b. The alternate names of FeS and Fe_2O_3 are ferrous sulfide and ferric oxide, respectively.

Sample Exercise 2.8 (cont.)

- We use the Roman numeral system for designating the charges on transition metal ions, but you may encounter *–ous* / *-ic* nomenclature in older books and articles.

Sample Exercise 2.9

Write the chemical formulas of (a) sodium sulfate and (b) magnesium phosphate.

- We are given the names of two compounds containing oxoanions and are to write their chemical formulas. The cations in these compounds are those formed by Na and Mg atoms.

Sample Exercise 2.9 (cont.)

- To write the formulas of these ionic compounds, we need to know the formulas and charges of the ions. Sodium is in group 1, and magnesium is in group 2. The charges on their ions are 1+ and 2+, respectively. The sulfate ion is SO_4^{2-}, and phosphate is $P°O_4^{3-}$.

a. To balance the charges on Na^+ and SO_4^2, we need twice as many Na^+ ions as SO_4^{2-} ions. Therefore the formula is Na_2SO_4.
b. $Mg_3(PO_4)_2$.

Sample Exercise 2.9 (cont.)

- To complete this exercise we had to know the formulas and charges of the sulfate and phosphate oxoanions. The charges on the cations could be inferred from the positions of the elements in the periodic table. In writing the formula, we used parentheses around the phosphate ion in magnesium phosphate to make it clear that the subscript 2 applies to the entire oxoanion.

Sample Exercise 2.10

Name the following compounds: (a) $CaCO_3$, (b) $LiNO_3$, (c) $MgSO_3$, (d) $RbNO_2$, (e) $KClO_3$, and (f) $NaHCO_3$.

- We are to name six compounds each containing an oxoanion. The names of ionic compounds begin with the names of the parent elements of the cations followed by the names of the oxoanions.

Sample Exercise 2.10 (cont.)

- The cations in these compounds are those formed by atoms of the elements (a) calcium, (b) lithium, (c) magnesium, (d) rubidium, (e) potassium, and (f) sodium. The names of the oxoanions, as listed in the book, are (a) carbonate, (b) nitrate, (c) sulfite, (d) nitrite, (e) chlorate, and (f) hydrogen carbonate.

- Combining the names of these cations and oxoanions, we get (a) calcium carbonate, (b) lithium nitrate, (c) magnesium sulfite, (d) rubidium nitrite, (e) potassium chlorate, and (f) sodium hydrogen carbonate.

Sample Exercise 2.10 (cont.)

- Sodium hydrogen carbonate is often called sodium bicarbonate. The prefix *bi-* is sometimes used to indicate that there is a hydrogen ion (H^+) attached to an oxoanion.

Sample Exercise 2.11

Name the oxoacids formed by the following oxoanions: (a) SO_3^{2-} ; (b) ClO_4^- ; (c) NO_3^-.

- We are given the formulas of three oxoanions and areto name the oxoacids formed when they combine with H^+ ions. Analyze According to Tables, the names of the oxoanions are (a) sulfite, (b) perchlorate, and (c) nitrate. When the oxoanion name ends in -*ite*, the corresponding oxoacid name ends in -*ous*. When the anion name ends in -*ate*, the oxoacid name ends in -*ic*.

Sample Exercise 2.11

- According to Tables, the names of the oxoanions are (a) sulfite, (b) perchlorate, and (c) nitrate. When the oxoanion name ends in *-ite*, the corresponding oxoacid name ends in *-ous*. When the anion name ends in *-ate*, the oxoacid name ends in *-ic*.

Sample Exercise 2.11 (cont.)

- Making the appropriate changes to the endings of the oxoanion names and adding the word *acid*, we get (a) sulfurous acid, (b) perchloric acid, and (c) nitric acid.
- Once we know the names of the common oxoanions, naming the corresponding oxoacids is simply a matter of changing the ending of the oxoanion name from *-ate* to *-ic*, or from *-ite* to *-ous*, and then adding the word *acid.*