CHEMISTRY

an atoms-focused approach

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Chapter 2

Atoms, Ions, and Molecules

The Building Blocks of Matter

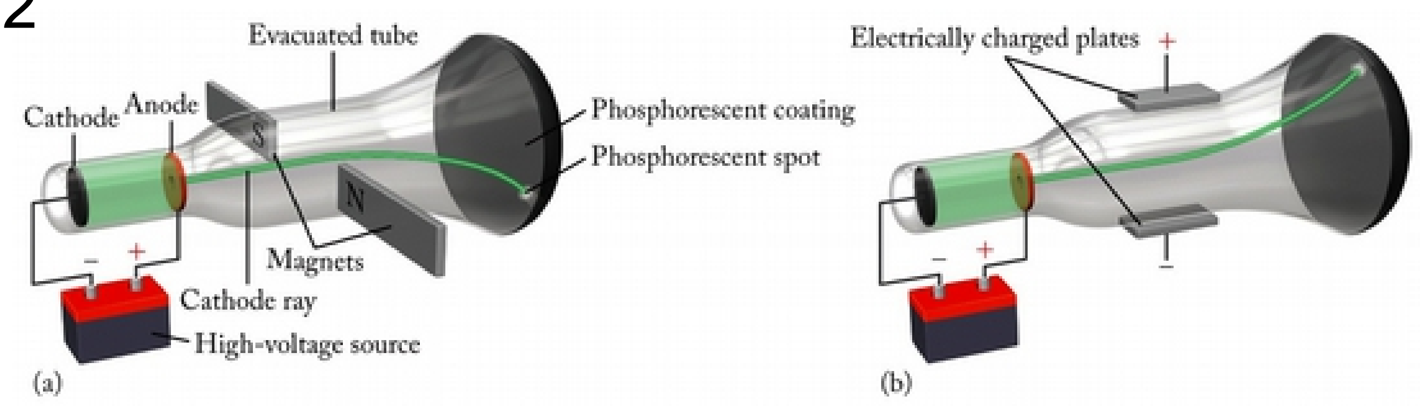
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides

Electrons

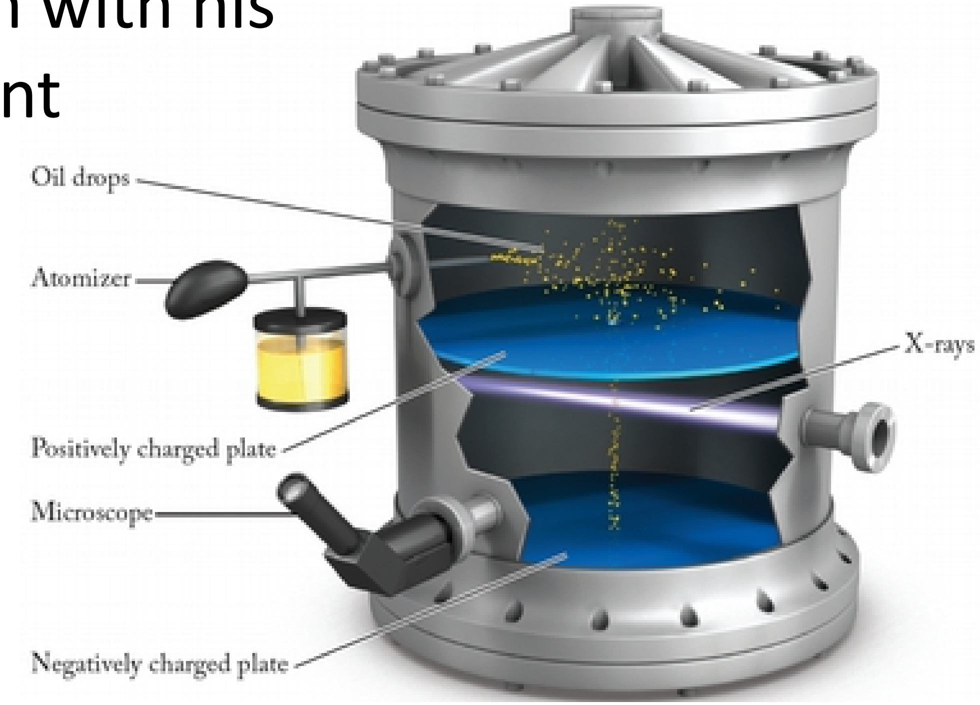
- J. J. Thomson (1897)
  - Beam from cathode ray tube was deflected toward positively charged plate
  - Atoms contain negatively charged particles with a constant mass-to-charge ratio

Figure 2.2



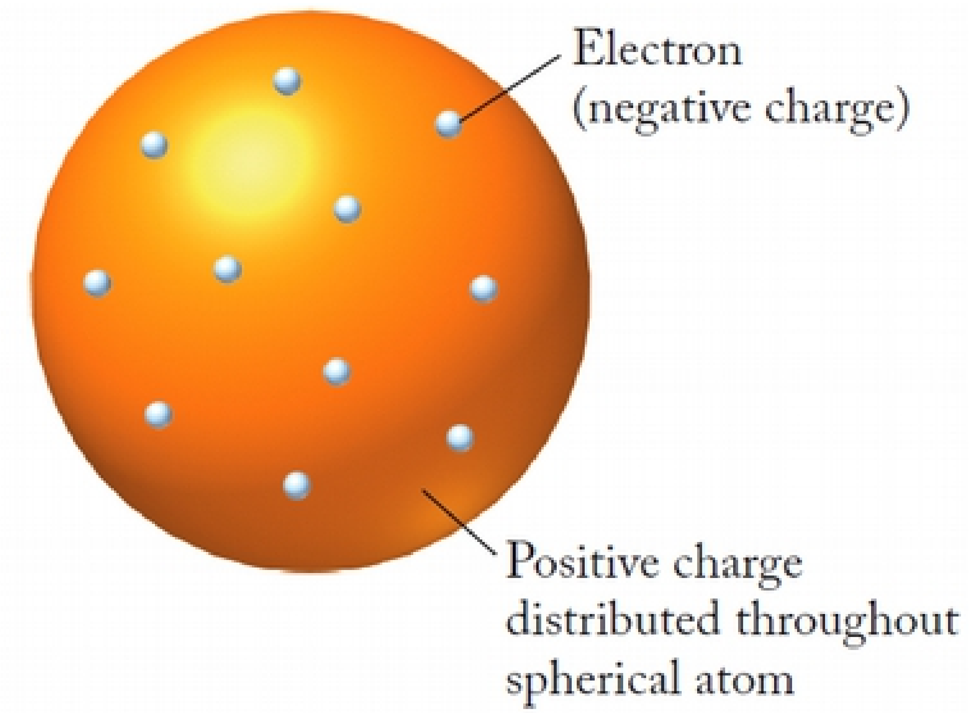
Electrons

- Robert Millikan (1909)
  - Determined the mass and charge of an electron with his oil-droplet experiment
  - $e^{-} = -1.602 \times 10^{-19}$ C
  - $m_e = 9.109 \times 10^{-28}$ g



Thomson's Plum-Pudding Model

- Plum-Pudding Model:
  - *e*⁻ distributed throughout diffuse, positively charged sphere

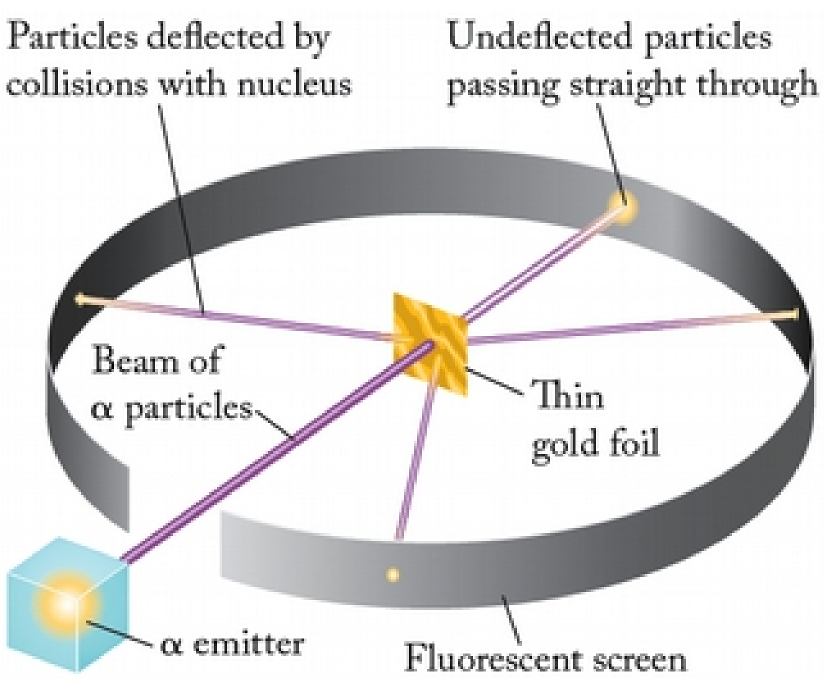


Radioactivity

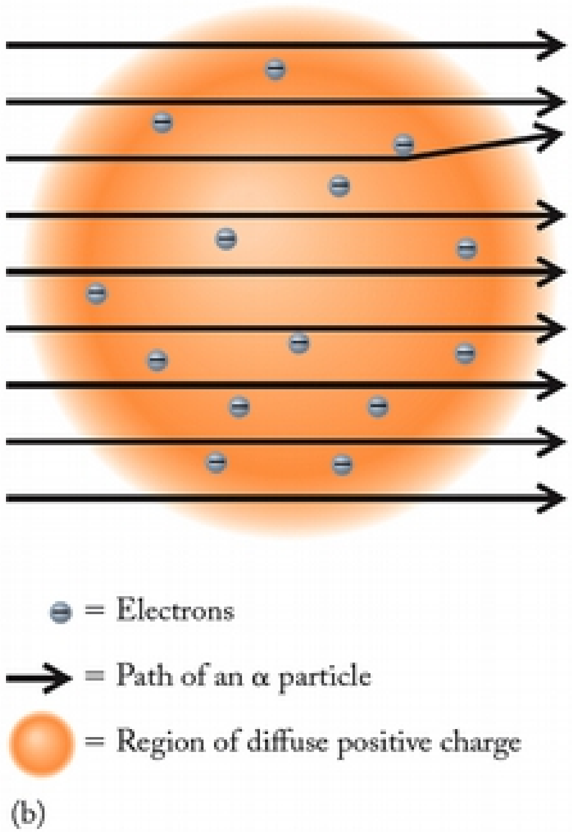
- Henri Becquerel (1896)
  - Some materials produce invisible radiation, consisting of charged particles
- Radioactivity
  - Spontaneous emission of high-energy radiation and particles
    - Beta particles (β, high-energy electrons)
    - Alpha particles (α, +2 charge, mass = He nucleus)

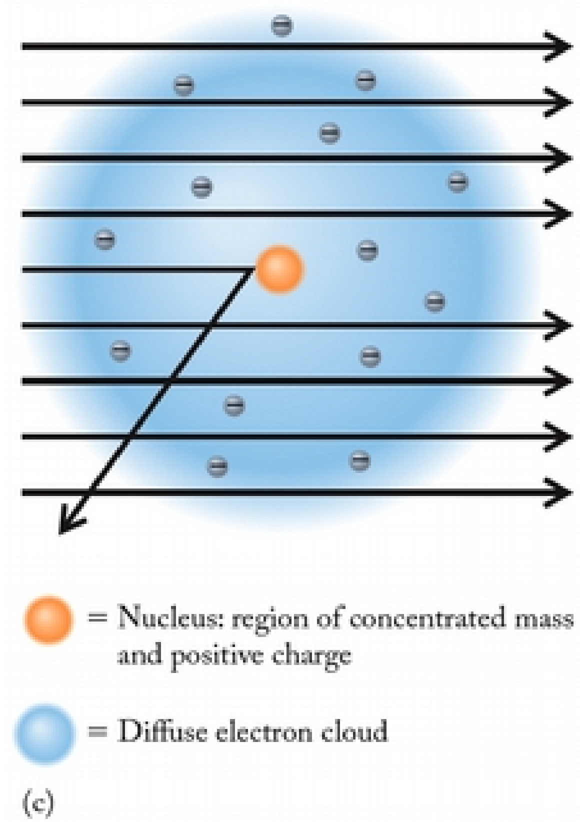
Rutherford’s Nuclear Model

- Rutherford’s Experiment:
  - Bombarded a thin gold foil with α particles to test Thomson’s model of the atom



Rutherford's Experiment





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

c) Actual results.

The Nuclear Atom

- The Nucleus:
  - Positively charged center of an atom, containing nearly all of the atom's mass
  - About 1/10,000 the size of the atom
  - Consists of two types of particles
    - Proton: Positively charged subatomic particle
      - Number defines the element
    - Neutron: Electrically neutral subatomic particle

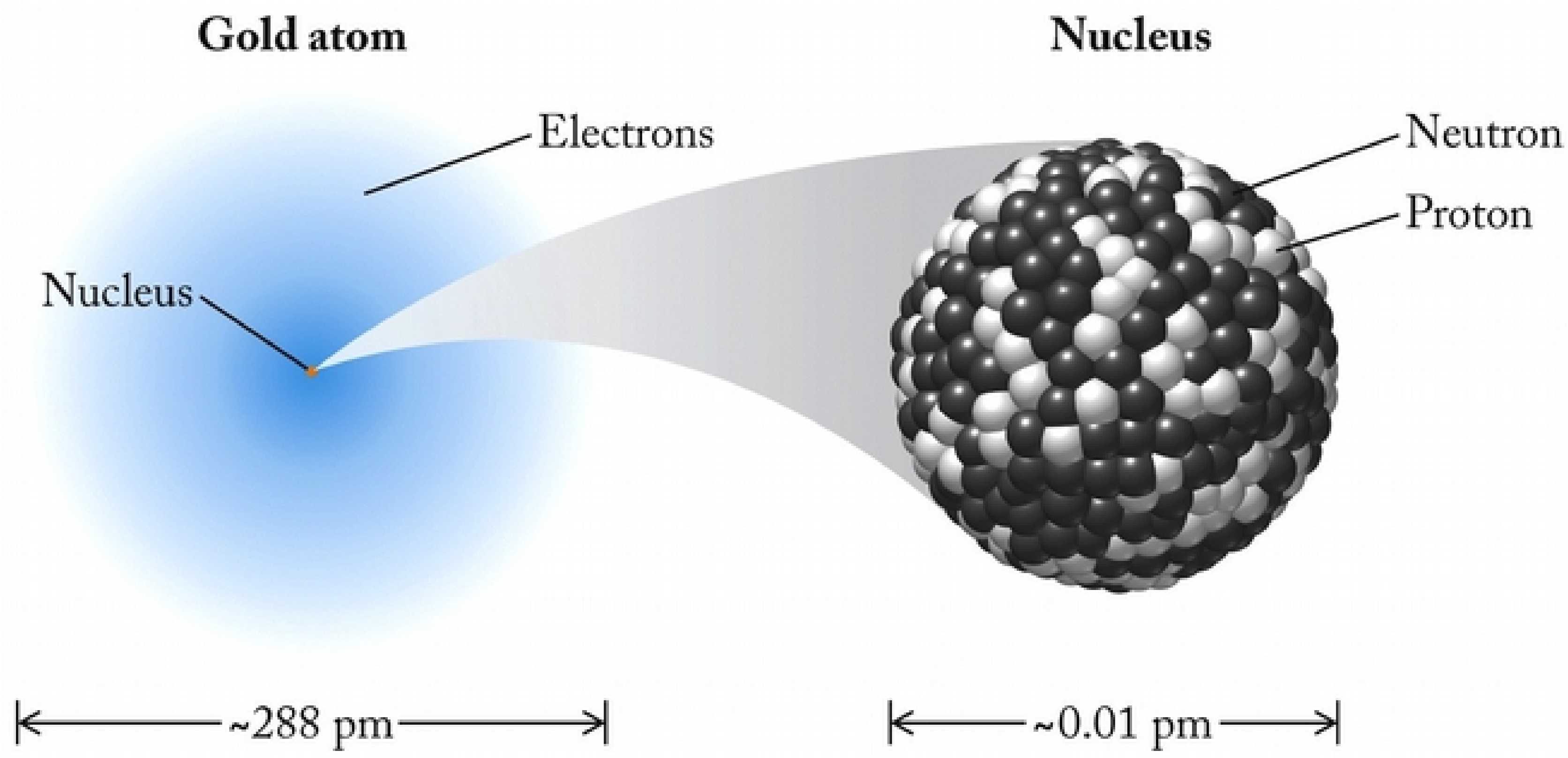
The Nuclear Atom

- Electrons:
  - Negatively charged particles
  - Roughly 2000 times smaller mass than protons
  - Located outside of the nucleus in orbitals or “electron clouds”
  - Outer electrons define the radius of the atom.
  - Electrons, and the nucleus, are much smaller than the atom itself, so most of the atom is empty space

The Nuclear Atom

- Ions:
  - Atoms are electrically neutral
    - This means that they have the same number of electrons as protons
  - Ions are formed when atoms gain or lose electrons (which are negatively charged)
    - Cations have lost electrons, so have a positive charge
    - Anions have gained electrons, so have a negative charge

The Nuclear Atom



Nucleus: Protons (+ charge) plus neutrons (neutral)

Atomic Mass Units

- Atomic Mass Units (amu)
  - Unit used to express the relative masses of atoms and subatomic particles
  - Equal to 1/12 of a carbon atom:
    - 6 protons + 6 neutrons = 12 amu
  - 1 amu = 1 dalton (Da)

Subatomic Particles

TABLE 2.1 Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Mass Number	Mass (kg)	Charge (relative value)	Charge (C)
Neutron	${}^{1}_{0}n$	1.00867	1	1.67493×10^{-27}	0	0
Proton	${}^{1}_{1}p$	1.00728	1	1.67262×10^{-27}	1+	$+1.602 \times 10^{-19}$
Electron	${}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$	5.485799×10^{-4}	0	9.10938×10^{-31}	1−	-1.602×10^{-19}

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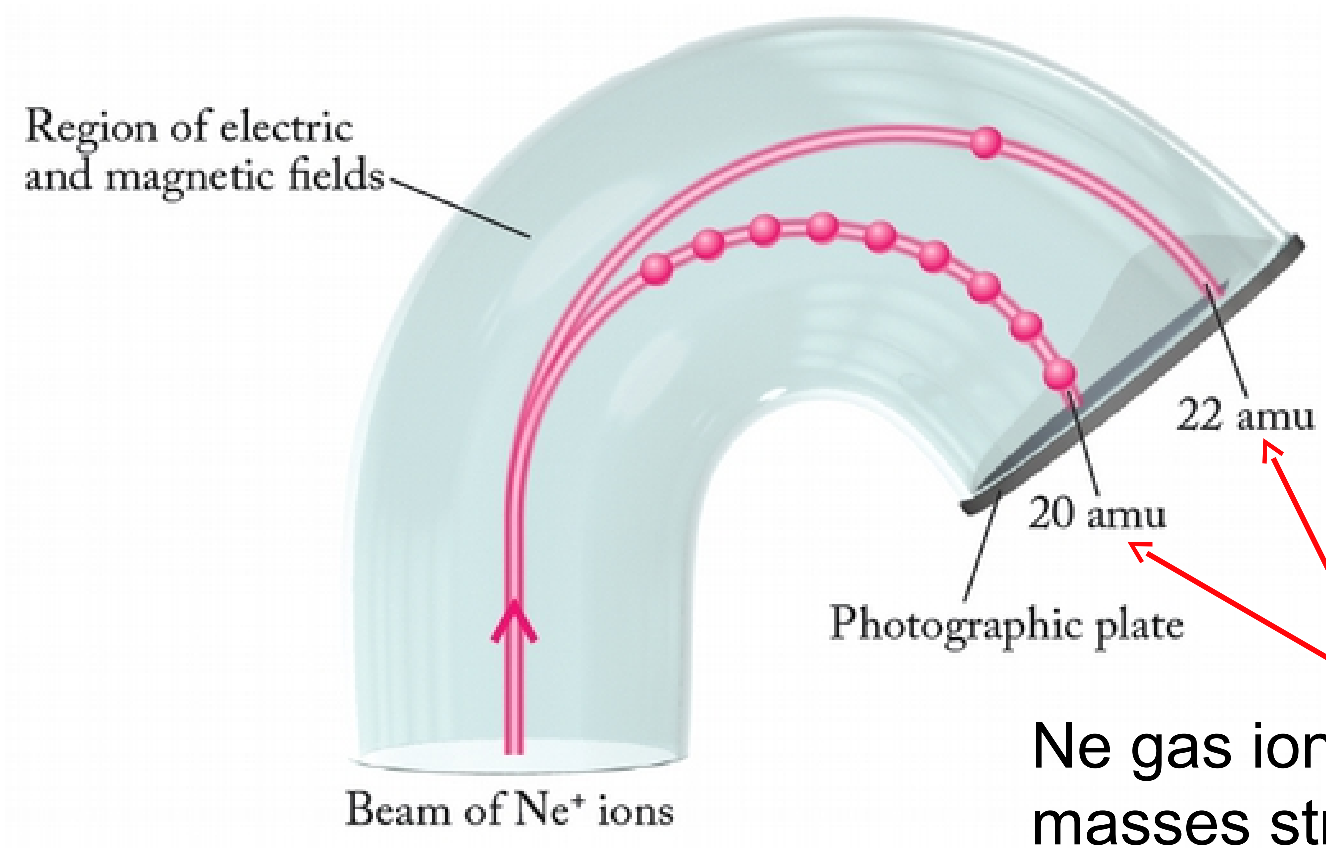
Isotopes: Experimental Evidence



Ne gas ions of different masses strike detector in different locations

Isotopes

- 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 an element containing the same # of protons but different # of neutrons
- Nuclide: A specific isotope of an element

Symbols of Nuclides

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

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

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

Symbols of Nuclides

- Z = atomic # = # of protons = p
- A = mass # = p + n
- Isotopes are denoted using the chemical symbol, X, or the element name:

  $_{Z}^{A}X$ = ^{A}X = X-A = element name-A

  So a carbon isotope with 6 neutrons could be written as:

  $_{6}^{12}C$ = ^{12}C = C-12 = Carbon-12

Practice: Isotopic Symbols

- Use the format ^{A}X to write the symbol for the nuclides having 28 protons and 31 neutrons

  - Collect and Organize:
  - Analyze:
  - Solve:
  - Think about It:

Practice: Identifying Atoms and Ions

- Complete the missing information in the table.

Symbol	^{23}Na	?	?	?
Number of Protons	?	39	?	79
Number of Neutrons	?	50	?	?
Number of Electrons	?	?	50	?
Mass Number	?	?	118	197

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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Mendeleev's Periodic Table

Dmitri 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

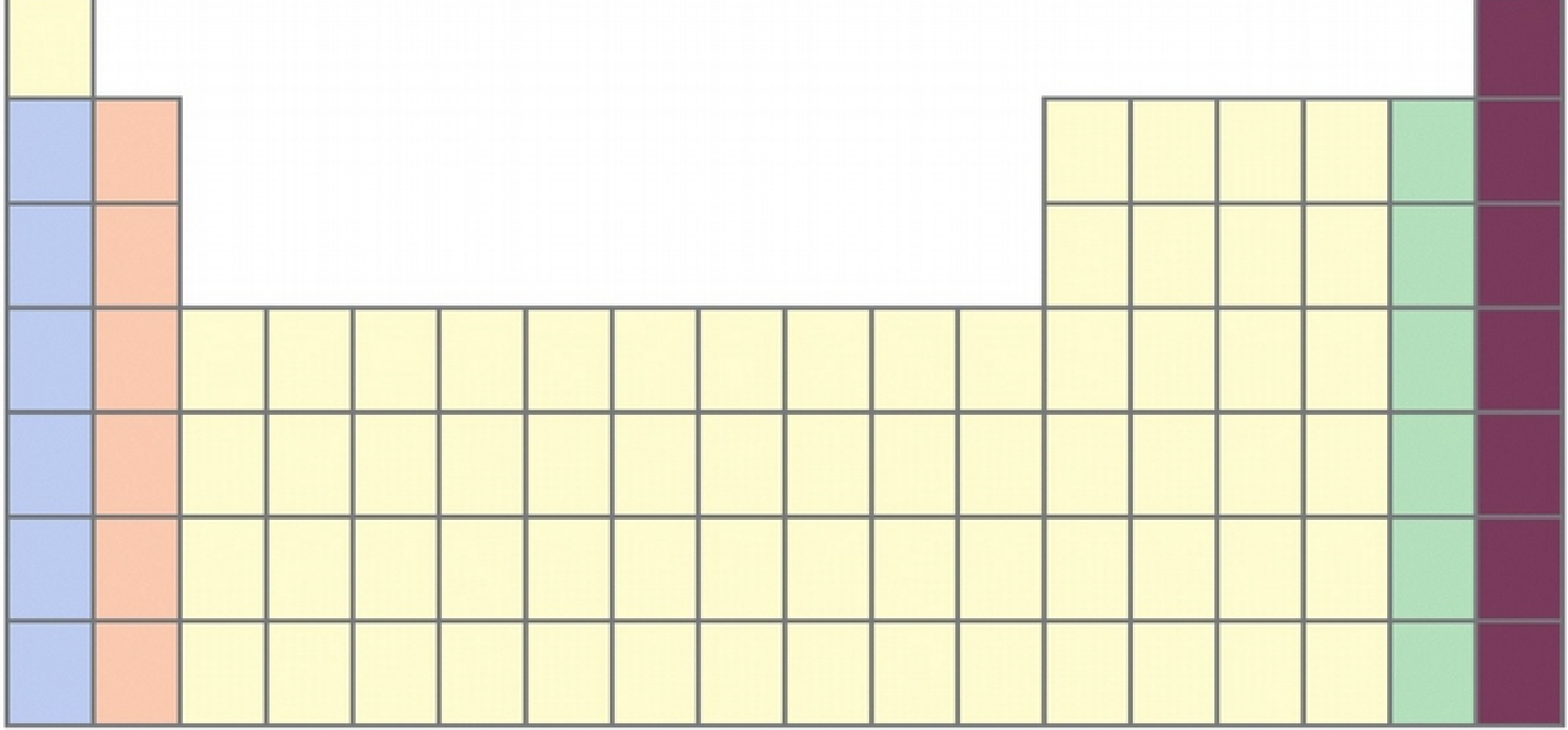
Group Number →

Row	I	II	III	IV	V	VI	VII	VIII
1	1 H							
2	7 Li	9.4 Be	11 B	12 C	14 N	16 O	19 F	
3	23 Na	24 Mg	27.3 Al	28 Si	31 P	32 S	35.5 Cl	
4	39 K	40 Ca	44 ?	48 Ti	51 V	52 Cr	55 Mn	56 Fe, 59 Co, 59 Ni
5	63 Cu	65 Zn	68 ?	72 ?	75 As	78 Se	80 Br	
6	85 Rb	87 Sr	88 ?Yt	90 Zr	94 Nb	96 Mo	100 ?	104 Ru, 104 Rh, 106 Pd
7	108 Ag	112 Cd	113 In	118 Sn	122 Sb	125 Te	127 J	
8	133 Cs	137 Ba	138 ?Di	140 ?Ce				
9								
10			178 ?Er	180 ?La	182 Ta	184 W		195 Os, 197 Ir, 198 Pt
11	199 Au	200 Hg	204 Tl	207 Pb	208 Bi			
12				231 Th		240 U		

Figure 2.9

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.



- Group 1: Alkali metals
- Group 2: Alkaline earth metals
- Group 17: Halogens
- Group 18: Noble gases

4 — Atomic number
Be — Symbol for element

Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

6 Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
7 Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

24 Cr 52.00	
24	Atomic number
52.00	Atomic mass

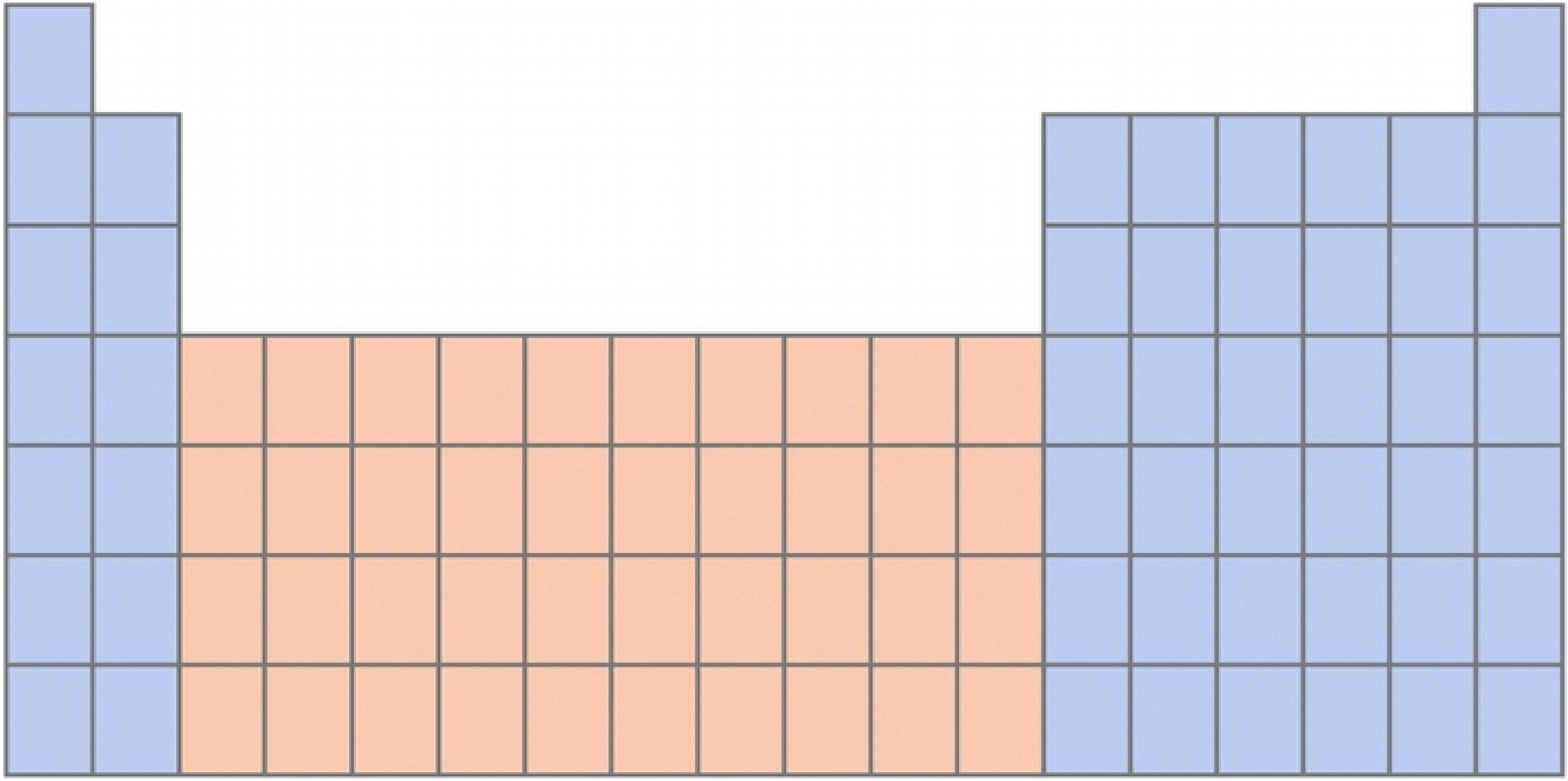
1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A
1 **H** 1.008																
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.61	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9
55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (210)	85 **At** (210)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (269)	109 **Mt** (268)	110	111	112	(113)	(114)	(115)	(116)	(117)

58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0
90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)

Metals

Metalloids

Nonmetals



Main group elements (representative elements)

Transition metals

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 (brittle), liquids and gases; nonconductors
- Metalloids (between metals/nonmetals)
  - Shiny solids (like metals); brittle (like nonmetals); semiconductors

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Average Atomic Mass

Most of the mass of an atom is in the nucleus, but the actual atomic mass is not exactly the sum of the masses of the nucleons (p + n).

- The electrons do have a small mass.
- Some mass is lost in the energy binding nucleons together ($E = m c^2$)

The periodic table shows the average atomic masses, in amu.

These masses are the weighted averages of the masses of all of the naturally occurring isotopes.

Average Atomic Mass

- Average Atomic Mass:
  - Weighted average of masses of all isotopes of an element
  - Calculated by multiplying the natural percent abundance of each isotope by its mass in amu and then summing these products
- Natural Abundance:
  - Proportion of a particular isotope, usually expressed as a percentage, relative to all the isotopes for that element in a natural sample
  - Assumes the same percentages over the surface of the Earth

Average Atomic Mass

Average atomic mass = sum over all isotopes of the mass of each isotope times its fractional abundance (percentage/100):

$$\text{Avg. at. mass} = \Sigma \frac{(\textit{atomic mass})(\textit{percent abundance})}{100\%}$$

where "Σ" means to sum over all isotopes present

Weighted Average

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:

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

Weighted Average

Lithium has two naturally occurring isotopes:
Li-6 6.015 amu   7.42 % abundance
Li-7 7.016 amu   92.58% abundance

So, for a standard sample of lithium:

avg. at. mass =

$$\frac{(6.015\ \text{amu})(7.42\%)}{100\%} + \frac{(7.016\ \text{amu})(92.58\%)}{100\%}$$

= 0.446 amu + 6.50 amu
= 6.94 amu

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The Mole

- A"mole" is a unit for a specific number:
  - 1 dozen items = 12 items
  - 1 mole particles = 6.022×10^{23} particles (also known as Avogadro's number)
- Exactly 12 grams of carbon-12 will contain exactly 1 mole of atoms
- Convenient unit for expressing macroscopic quantities (atoms or molecules) involved in macroscopic processes we observe

Moles as Conversion Factor

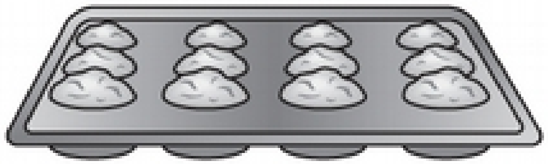
- To convert between number of particles and an equivalent number of moles

$$\text{Number of } \cancel{\text{particles}} \times \frac{1 \text{ mole}}{6.022 \times 10^{23}\ \cancel{\text{particles}}} = \text{Number of moles}$$

$$\text{Number of } \cancel{\text{moles}} \times \frac{6.022 \times 10^{23} \text{ particles}}{1\ \cancel{\text{mole}}} = \text{Number of particles}$$

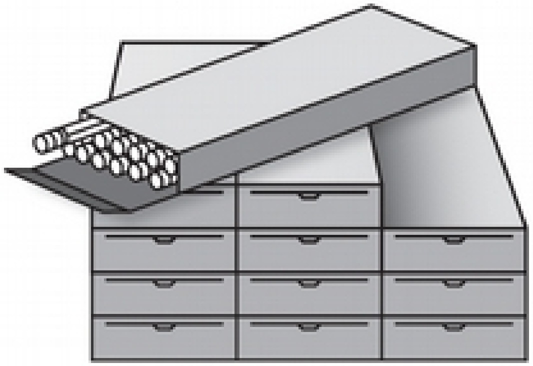
2 gloves—1 pair of gloves

(a)



12 rolls—1 dozen rolls

(b)

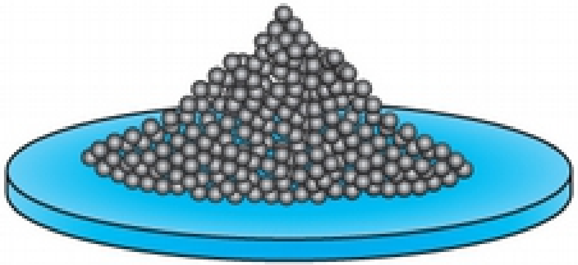


144 pencils—1 gross of pencils

(c)

500 sheets of paper—1 ream of paper

(d)



6.022×10^{23} iron atoms—1 mole of iron atoms

(e)

Moles as Conversion Factor

***To convert between moles and numbers*:**

Use the conversion factor:

1 mole = 6.022E23

For example:

$$1.204\text{E}24\ atoms\left(\frac{1\ mole\ atoms}{6.022\text{E}23\ atoms}\right)=2.000\ moles\ atoms$$

$$2.50\ moles\ atoms\left(\frac{6.022\text{E}23\ atoms}{1\ mole\ atoms}\right)=1.506\text{E}24\ atoms$$

Moles as Conversion Factor

$$1.5\,mole\,molecules\left(\frac{6.022\text{E}23\,molecules}{1\,mole\,molecules}\right)=9.03\text{E}23\,molecules$$

$$1.5\,mole\,H_2O\left(\frac{6.022\text{E}23\,H_2O}{1\,mole\,H_2O}\right)=9.03\text{E}23\,H_2O\,(molecules)$$

$$3.01\text{E}23\,H_2O\left(\frac{1\,mole\,H_2O}{6.022\text{E}23\,H_2O}\right)=0.500\,mole\,H_2O$$

Molar Mass

- Molar Mass:
  - The mass (in grams) of one mole of the substance (atom, molecule or formula unit)
    - 1 atom of He = 4.003 amu
    - Mass of 1 mole of He atoms = 4.003 g
  - The molar mass ($\mathcal{M}$) of He is 4.003 g/mol

Molar Mass

- Molar Mass:
  - The mass (in grams) of one mole of the substance (atom, molecule or formula unit)
- The molar mass ($\mathcal{M}$) of He is 4.003 g/mol
- The average mass on PT gives:
  - mass, in amu, of one atom of the element
    - 1 atom of He = 4.003 amu He
  - mass, in grams, of one mole of atoms of the element
    - 1 mole of He atoms = 4.003 g He

Molar Mass

To find the moles of hydrogen atoms in 1.01 grams of H:

$$1.01\ g\ H\left(\frac{1\ mole\ H}{1.01\ g\ H}\right)=1.00\ mole\ H$$

or for 12.5 grams of hydrogen atoms:

$$12.5\ g\ H\left(\frac{1\ mole\ H}{1.01\ g\ H}\right)=12.4\ mole\ H$$

Molar Mass

Or to find the mass of 2.5 moles of helium:

$$2.5\, moles\ He\left(\frac{4.00\ g\ He}{1\, mole\ He}\right) = 10.\ g\ He$$

or for 2.5 moles of lead:

$$2.5\, moles\ Pb\left(\frac{207.20\ g\ Pb}{1\, mole\ Pb}\right) = 518.0\ g\ Pb$$

Note these two examples have the same number of atoms, but different masses

Molecular Mass/Formula Mass

- Molecular Mass:
  - Mass of one molecule of a molecular compound
  - Sum of the atomic masses of the atoms in that compound
  - Example: For one molecule of CO_2
    - CO_2 = C + 2 O
      = 12.01 amu + 2(16.00amu)
      = 44.01 amu/molecule
- Formula Mass:
  - Mass in atomic mass units of one formula unit of an ionic compound (e.g., NaCl)

Molecular Mass/Formula Mass

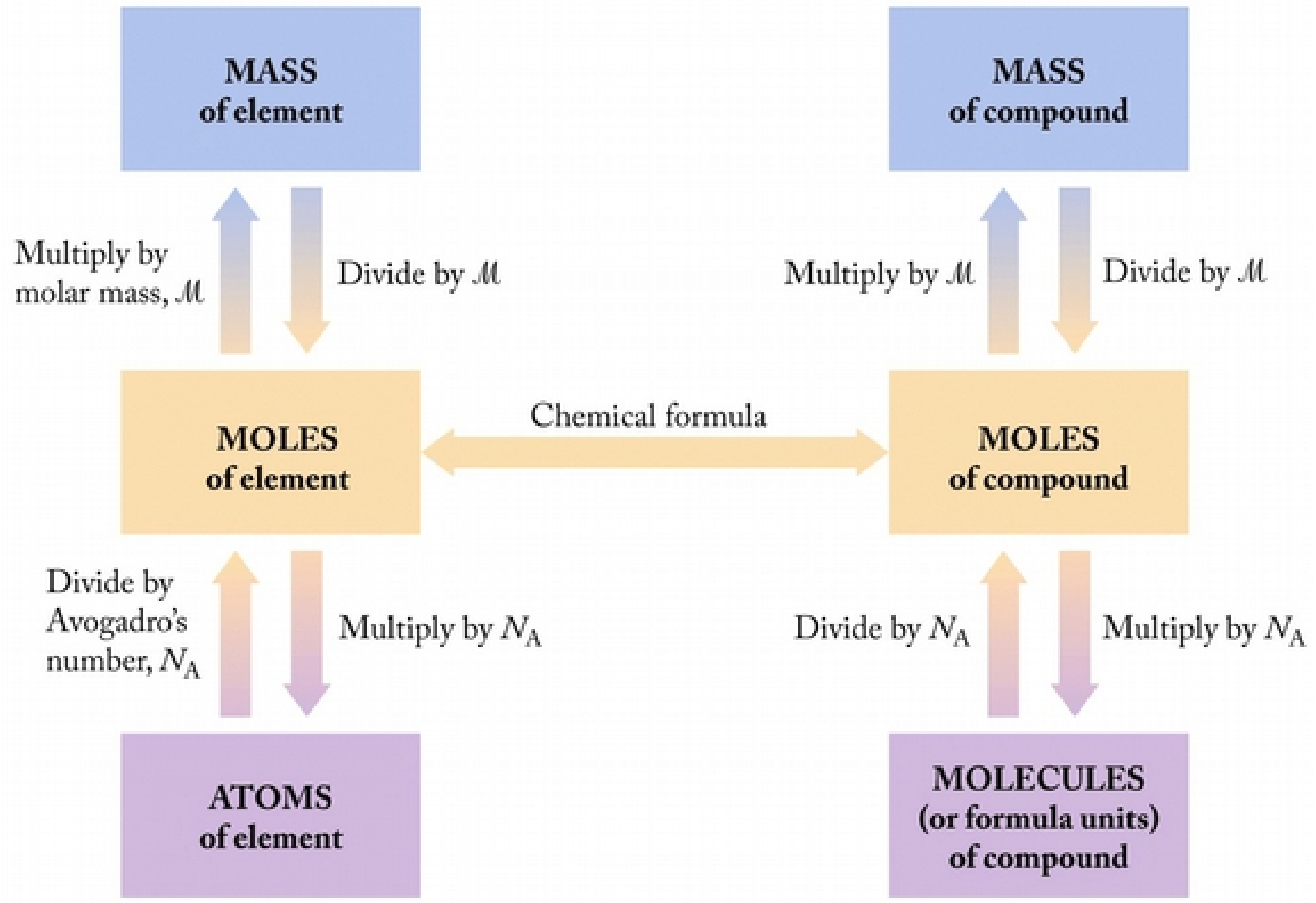
So, to find the mass of 1.5 moles of CO_2:

$$1.5\,moles\,CO_2\left(\frac{44.01\,g\,CO_2}{1\,mole\,CO_2}\right)=66.02\,g\,CO_2$$

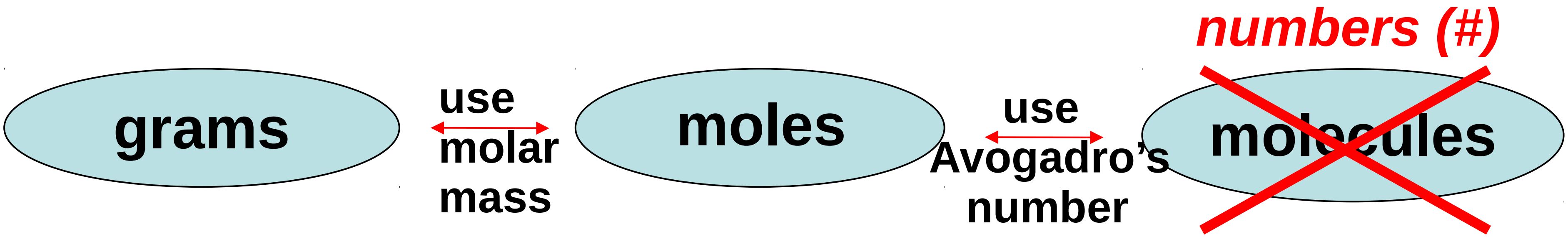
or to find the moles of molecules of 85.0 grams of CO_2:

$$85.0\,g\,CO_2\left(\frac{1\,mole\,CO_2}{44.01\,g\,CO_2}\right)=1.93\,mole\,CO_2$$

Moles, Mass, and Particles



Moles, Mass, and Particles



Practice: Mole Calculations #1

a) How many moles of K atoms are present in 19.5 g of potassium?

b) How many formula units are present in 5.32 moles of baking soda ($NaHCO_3$)?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Practice: Mole Calculations #2

How many moles are present in 58.4 g of chalk ($CaCO_3$)?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Practice: Mole Calculations #3

The uranium used in nuclear fuel exists in nature in several minerals. Calculate how many moles of uranium are found in 100.0 grams of carnotite, $K_2(UO_2)_2(VO_4)_2 \cdot 3H_2O$.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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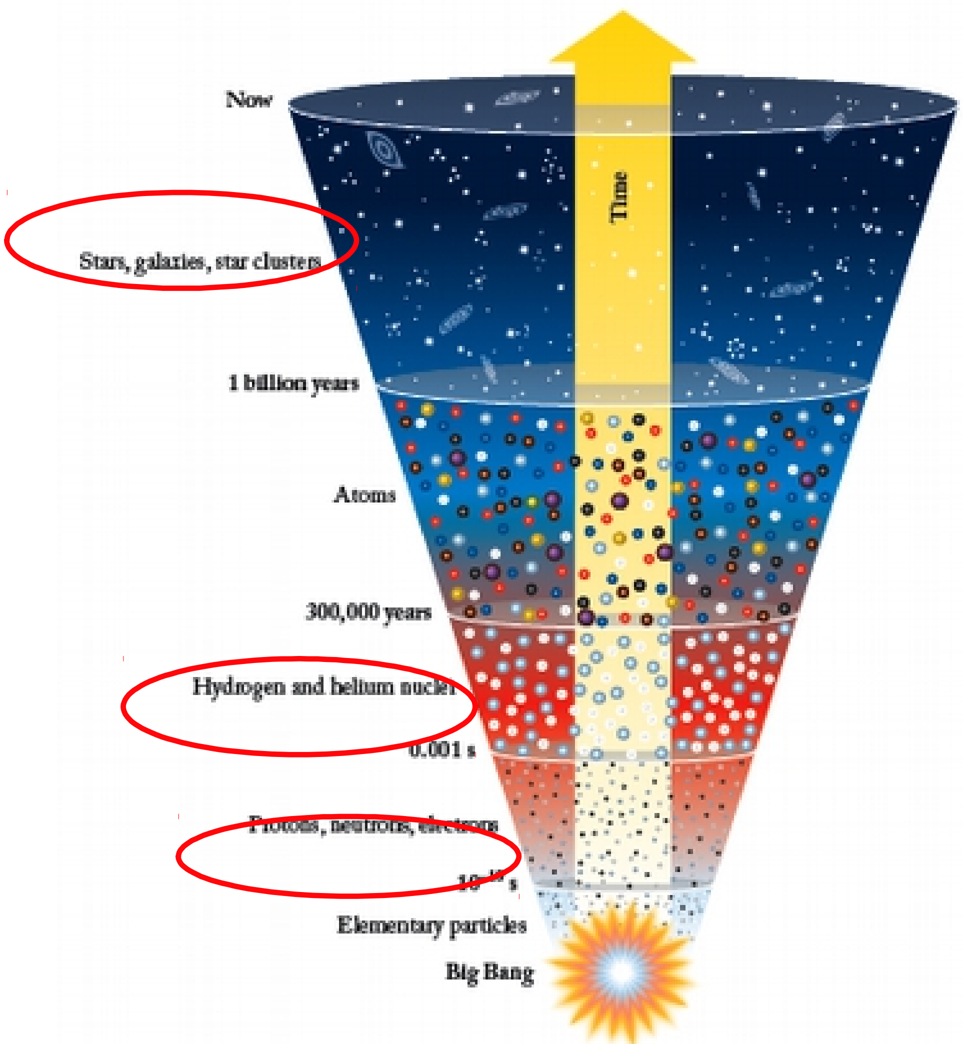
Big Bang Theory



H and He atoms in stars fuse to form heavier elements.

Subatomic particles fuse to form H and He nuclei.

Existence of subatomic particles.

Nucleosynthesis

- Nucleosynthesis:
  - Energy from Big Bang transformed into matter (more details of this matter/energy relationship in Chapter 21)
  - Fusing of fundamental/subatomic particles (protons/neutrons) created atomic nuclei

$$ {}^{1}_{1}\mathrm{p} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{2}_{1}\mathrm{d} $$

$$ 2\,{}^{2}_{1}\mathrm{d} \rightarrow {}^{4}_{2}\alpha $$

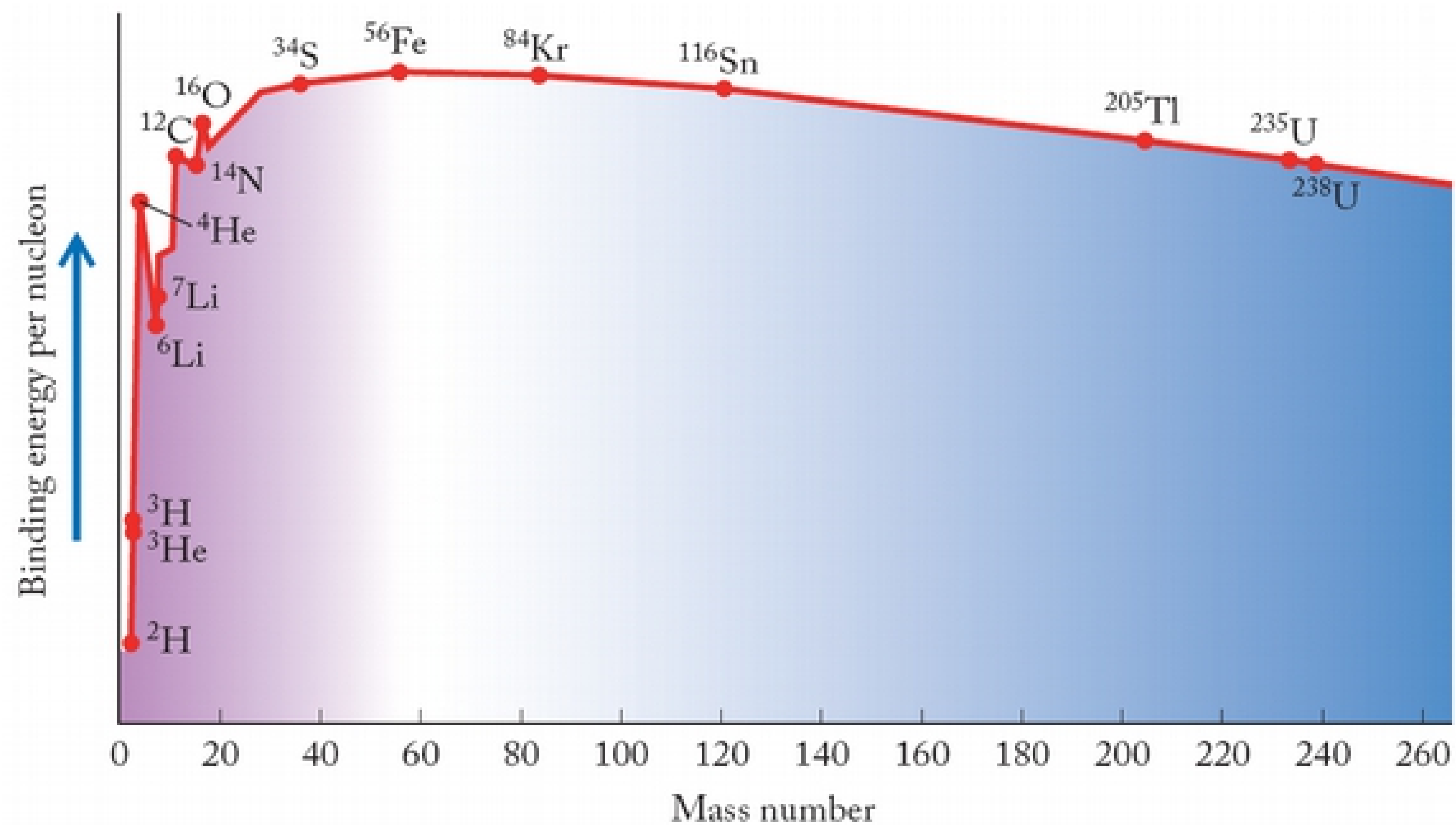
Nuclear Binding Energies

- The stability of a nucleus is proportional to its binding energy (BE)
  - $E = (\Delta m)c^2$
  - Δm = mass defect of the nucleus (in kg).
  - c = speed of light (2.998×10^8 m/s^2)

- Mass defect (Δm) – difference between the mass of the stable nucleus and the masses of the individual nucleons that compose it.

Stability of Nuclei

- Stability: Proportional to BE/# of nucleons
- ^{56}Fe = most stable nucleus



Stellar Nucleosynthesis

- High density and temperature in stars caused additional fusion reactions to create elements heavier than H, He:

$$3\ {}^{4}_{2}\alpha \rightarrow {}^{12}_{6}\mathrm{C}$$

$${}^{12}_{6}\mathrm{C} + {}^{4}_{2}\alpha \rightarrow {}^{16}_{8}\mathrm{O}$$

- Stellar core forms shells of heavier elements produced from fusion of lighter elements

Alpha (α) Decay

- Alpha Decay: Nuclear reaction in which an unstable nuclide spontaneously emits an alpha particle
  - α particle = He nucleus
- Example:

$$^{238}_{92}\text{U} \longrightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\alpha$$

Beta (β) Decay

- Beta Decay: Spontaneous ejection of a β-particle (electron) by a neutron-rich nucleus

$$ {}^{1}_{0}n \longrightarrow {}^{1}_{1}p + {}^{0}_{-1}e $$

- Example: ${}^{14}_{6}C \longrightarrow {}^{14}_{7}N + {}^{0}_{-1}\beta$

- (Note: mass and charge balance!)

Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides

Practice: Decay Equation

Radioactive radon-222 decays with a loss of one α particle. Write the balanced equation for this decay.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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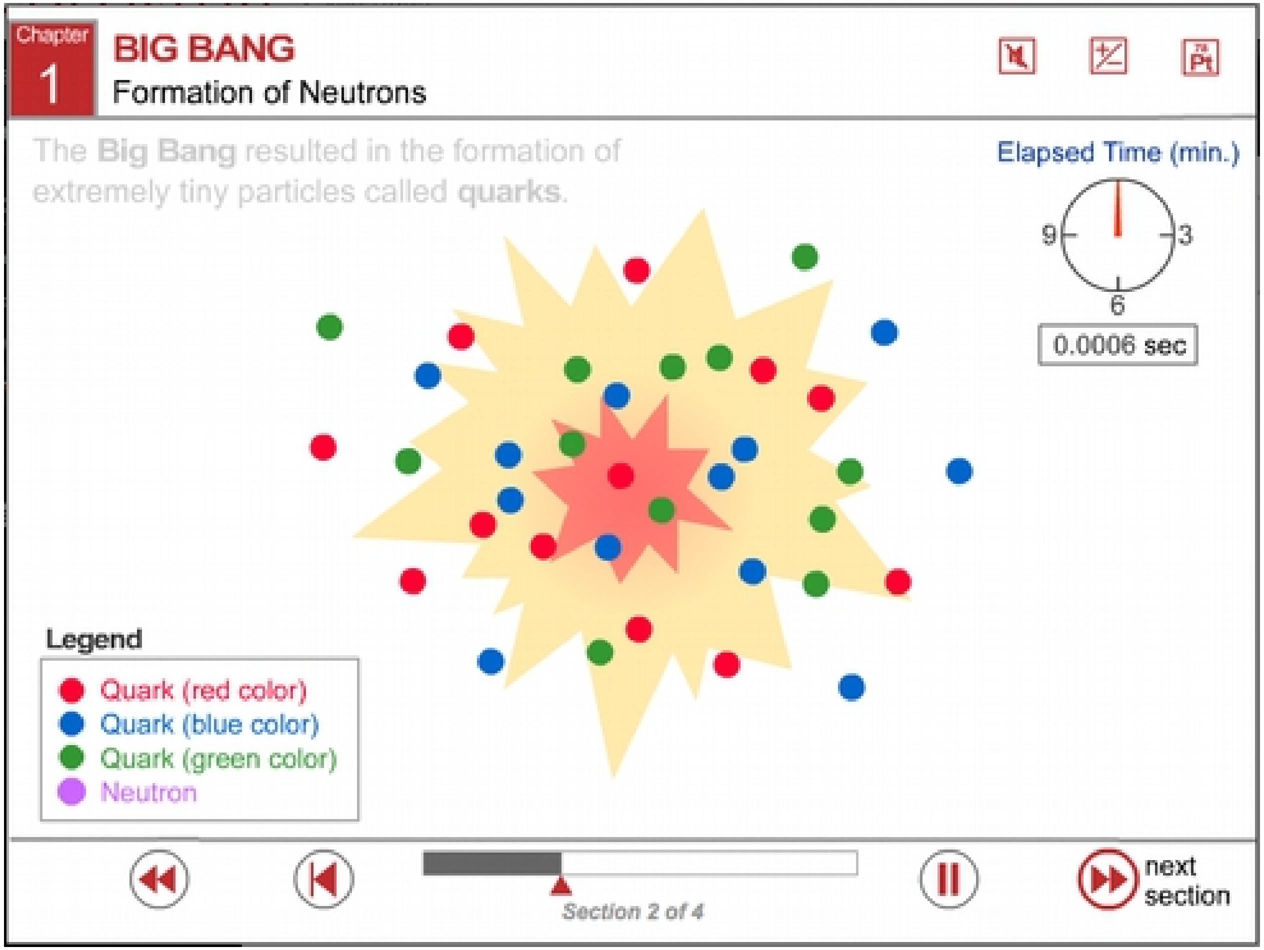

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This concludes the Lecture PowerPoint presentation for Chapter 2

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