Ch. 9 – Moles

Law of definite proportions – for a pure substance, each element is always present in the same proportion by mass.

- Also, for a pure substance, each element is always present in whole number ratios relative to the other elements

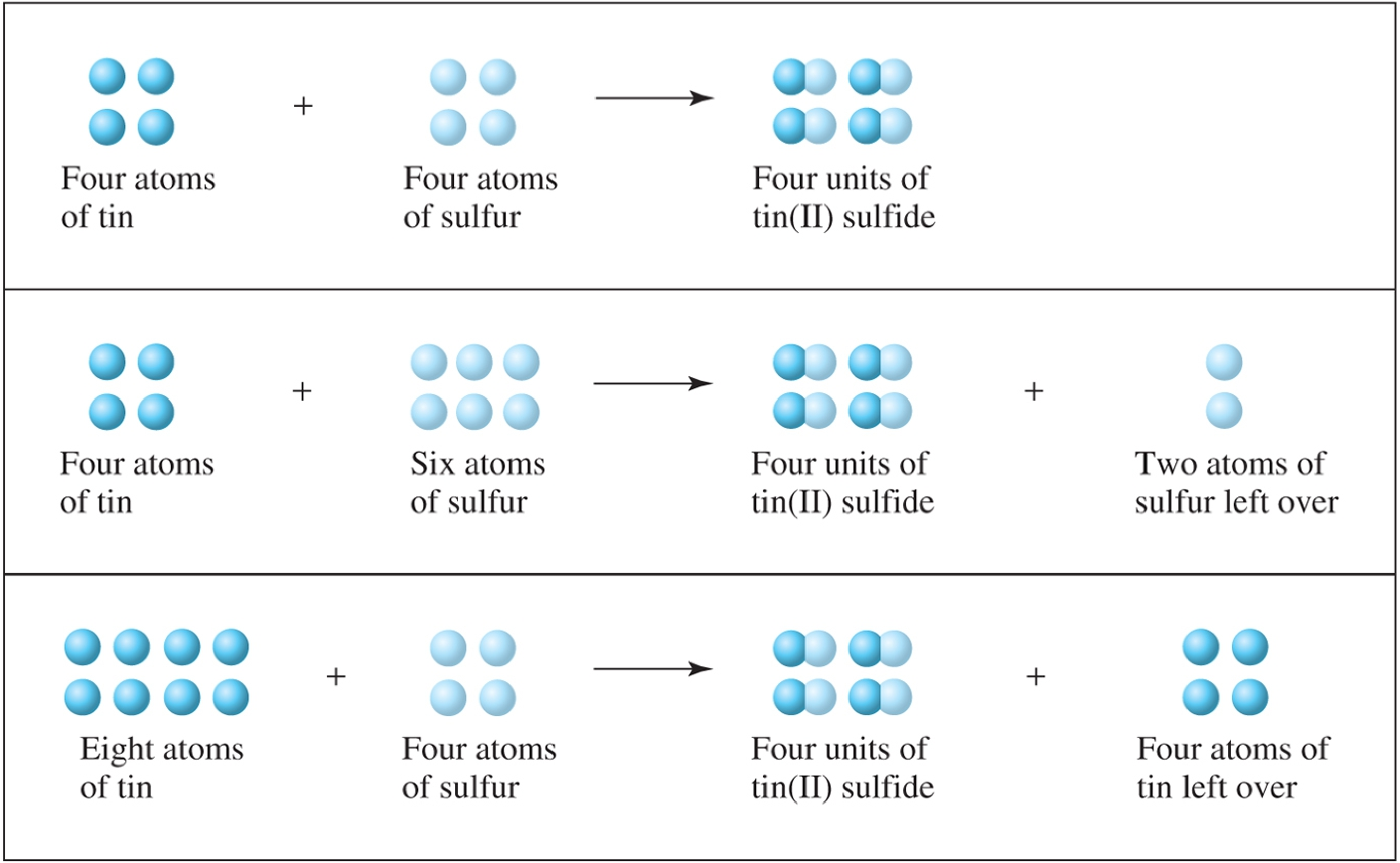

Four atoms of tin
+
Four atoms of sulfur
Four units of tin(II) sulfide
Four atoms of tin
+
Six atoms of sulfur
Four units of tin(II) sulfide
+
Two atoms of sulfur left over
Eight atoms of tin
+
Four atoms of sulfur
Four units of tin(II) sulfide
+
Four atoms of tin left over

TABLE 9.1 Data Illustrating the Law of Definite Proportions

Mass of Ca Used (g)	Mass of S Used (g)	Mass of CaS Formed (g)	Mass of Excess Unreacted Sulfur (g)	Ratio in Which Substances React
55.6	44.4	100.0	none	1.25
55.6	50.0	100.0	5.6	1.25
55.6	100.0	100.0	55.6	1.25
55.6	200.0	100.0	155.6	1.25
111.2	88.8	200.0	none	1.25

- Atomic mass – Actual mass of an atom
  - slightly different than mass # due to electrons and nuclear binding energy
- Average atomic mass
  - Most elements in nature are composed of multiple isotopes
  - The decimal value on the periodic table gives the weighted average of these isotopes in amu (atomic mass units)

Formula Mass

- The average mass of a compound
- Calculated by adding up the atomic masses of all of the atoms in one formula

For CO_2:

- 1 carbon atom + 2 oxygen atoms
- formula mass = 12.01 + 2(16.00)
  = 44.01 amu

Percent by mass

(% by mass of an element in a compound)

% by mass of X = (mass of X)/(total mass) * 100%

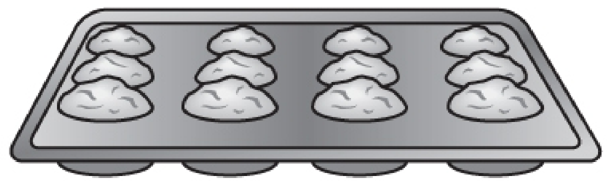
- If the atomic mass of an isotope is known, then any mass of that isotope can be converted to a number of atoms
- If the average mass of an element is known, then the number of atoms can also be calculated
- If the mass of a molecule is known, the we can also calculate the number of molecules in a sample from the mass

The Mole

- **The key to converting between mass and numbers is the mole**
- **'Mole' is just a number**
    - **1 mole (mol) = 6.022E23 (Avogadro's #)**
    - **We can have a mole of anything:**
        - **atoms**
        - **molecules**
        - **grains of sand....**

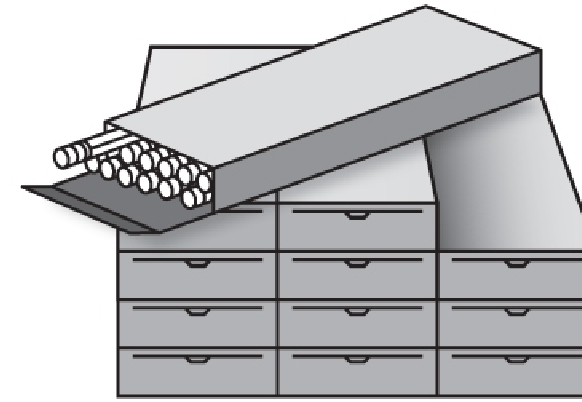
2 gloves—1 pair

(a)



12 rolls—1 dozen

(b)

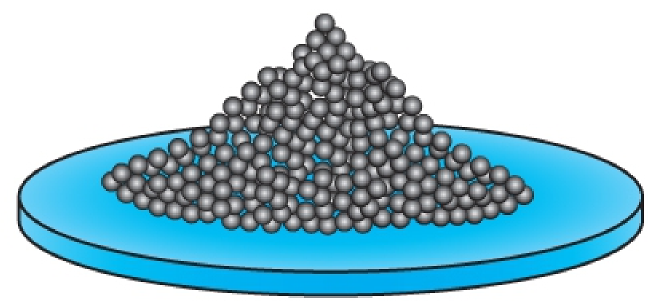


144 pencils—1 gross

(c)

500 sheets of paper—1 ream

(d)



6.022×10^{23} iron atoms—1 mole

(e)

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

Use the conversion factor:

1 mole = 6.022E23

For example:

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

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

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

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

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

- **One mole of protons (or neutrons) has a mass of one gram**
  - **(exactly one mole of C-12 atoms has a mass of exactly 12 grams)**
  - **The decimal value on the PT (the average atomic mass) may also be used to convert mass to moles**
    - **This value gives the mass, in grams, of one mole of the element**

Key	
24	Atomic number
Cr	
52.00	Atomic mass

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H** 1.008																	2 **He** 4.003
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	10 **Ne** 20.18
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	18 **Ar** 39.95
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	36 **Kr** 83.80
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	54 **Xe** 131.3
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)	86 **Rn** (222)
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)	(118)

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	71 **Lu** 175.0
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)	103 **Lr** (262)

Metals

Metalloids

Nonmetals

The 1–18 group designation has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) but is not yet in wide use. In this text we use the standard U.S. notation for group numbers (1A–8A and 1B–8B). No names have been assigned for elements 110–112. Elements 113–118 have not yet been synthesized.

Source: Raymond Chang, *General Chemistry: The Essential Concepts*, Third Edition, Copyright 2003 The McGraw-Hill Companies, New York, NY.

For example, to find the moles of hydrogen atoms in 1.01 grams:

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

or for 12.5 grams of hydrogen atoms:

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

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

Formula Mass

- **The average mass of a compound**
- **Calculated by adding up the atomic masses of all of the atoms in one formula**

For CO_2:

- **1 carbon atom + 2 oxygen atoms**
- **formula mass = 12.01 + 2(16.00)**
  = 44.01 amu
- **also: 1 mole CO_2 = 44.01 g CO_2**
- **molar mass = 44.01 g CO_2 / mol CO_2**

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

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

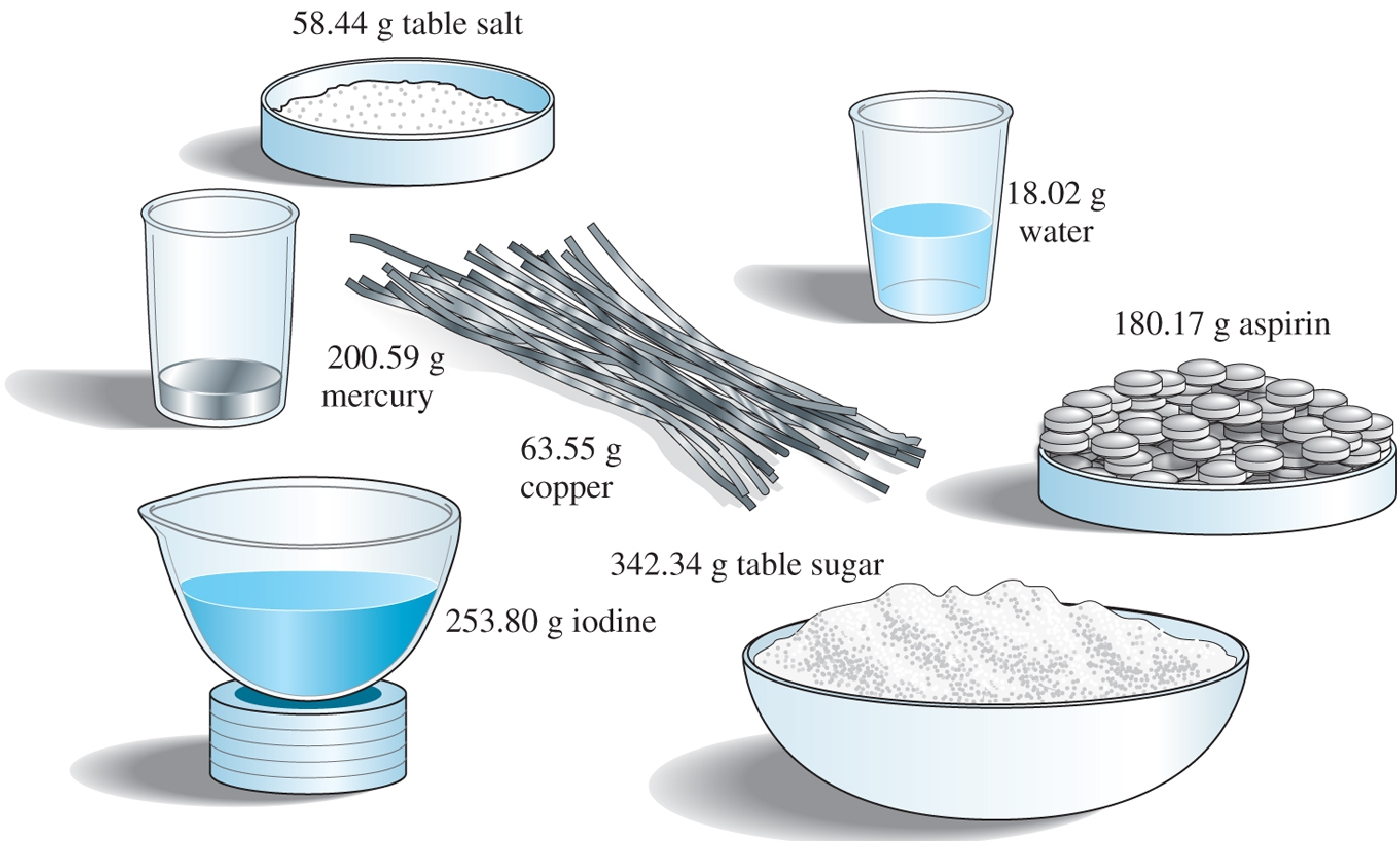
or to find the moles of 85.0 grams of CO_2:

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

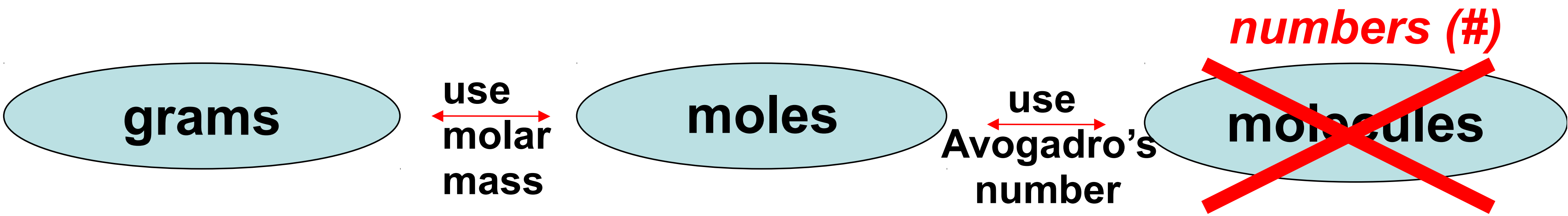
TABLE 9.2 Mole Relationships

Name	Formula	Formula Mass (amu)	Mass of 1 Mole Formula Units (g)	Number and Kind of Particles in 1 Mole
Atomic nitrogen	N	14.01	14.01	6.022×10^{23} N atoms
Molecular nitrogen	N_2	28.02	28.02	6.022×10^{23} N_2 molecules $2(6.022 \times 10^{23})$ N atoms
Zinc	Zn	65.38	65.38	6.022×10^{23} Zn atoms
Zinc ions	Zn^{2+}	65.38*	65.38	6.022×10^{23} Zn^{2+} ions
Calcium chloride	$CaCl_2$	110.98	110.98	6.022×10^{23} $CaCl_2$ units 6.022×10^{23} Ca^{2+} ions $2(6.022 \times 10^{23})$ Cl^- ions

**Recall that the electron has negligible mass; thus ions and atoms have essentially the same mass.*


58.44 g table salt
18.02 g water
200.59 g mercury
63.55 g copper
180.17 g aspirin
342.34 g table sugar
253.80 g iodine

Keep these relationships in mind:



Remember – the critical link between moles and grams of a substance is the molar mass.

IT'S SIMPLE – THINK IN TERMS OF PARTICLES!

Purity of Samples

$$\%\text{Purity} = \frac{mass\ of\ the\ compound}{total\ mass} * 100\%$$

To find the mass of a compound in an impure sample:

$$mass\ of\ the\ compound = \frac{\%\text{Purity}}{100\%} * \text{total mass}$$

Empirical Formula

- Empirical formula – subscripts show the lowest whole number ratios of elements

- May be determined from percent by mass data:
  - Assume a 100.00 g sample
  - Determine the mass of each element in the 100 g
  - Convert mass to moles
  - Determine the lowest whole number ratios

For ionic compounds, the empirical formula is the chemical formula.

For molecular compounds, more data is needed to determine the molecular formula

- The molar mass usually needs to be determined
- This molar mass will be a multiple of the molar mass of the empirical formula
- Divide the molecules molar mass by the empirical formulas molar mass
- Multiply the subscripts by this integer to get the molecular formula

TABLE 9.3 A Comparison of Empirical and Molecular Formulas for Selected Compounds

Compound	Empirical Formula	Molecular Formula	Whole-Number Multiplier
Dinitrogen tetrafluoride	NF_2	N_2F_4	2
Hydrogen peroxide	HO	H_2O_2	2
Sodium chloride	$NaCl$	$NaCl$	1
Benzene	CH	C_6H_6	6