Ch02

Molar Mass

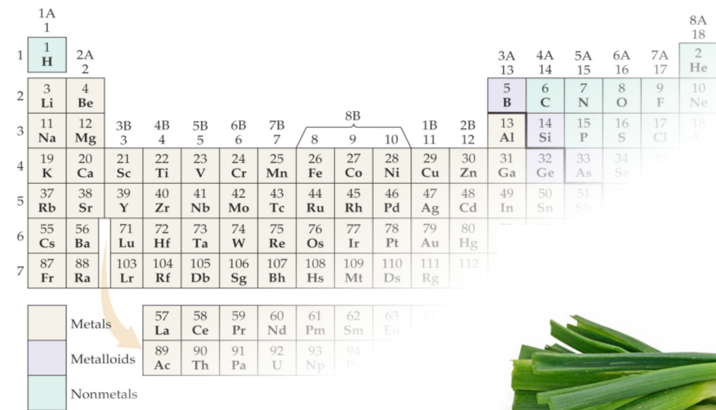
The weight of 6.022×10^{23} singles

The chemists dozen.

version 1.5

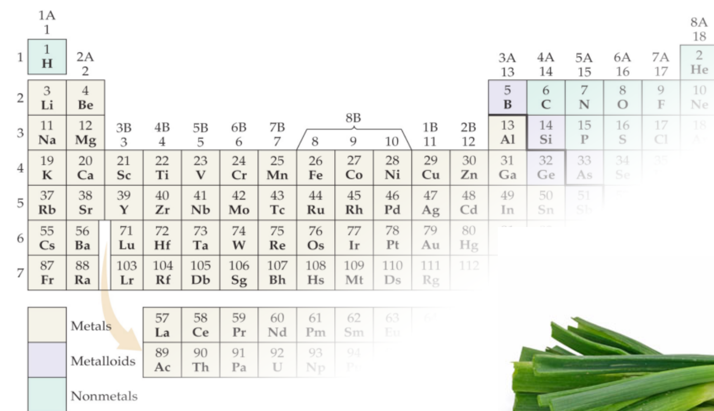
Molar Mass

- Counting by Weight
  - Bankers weigh coins, Chef's weigh scallions
  - Atoms are made of protons & neutrons
  - AMU unit is about the weight of a proton/neutron
  - We work in average weights
- The Chemists Dozen, the Mole
  - Defining the Mole
  - The Mole scales between amu and grams
    - calculations with mols
  - Mapping out more complicated problems
  - Four Essential Conversion Factors
    - Avogadro's Number
    - Molecular Formula
    - Formula Weight
      (aka Molecular Weight, Formula Mass)
    - Molar Weight (aka Molar Mass)
- Illustrative Problems
    - grams to atoms
    - molecules to grams



Counting by Weight

- A banker doesn't count pennies.
  - He know's how much a penny weighs. If you give him a bag of pennies he will weigh the bag, divide it by a pennies average weight and tell you the bags value.
- A banquet chef does the same.
  - If a recipe calls for 2 scallions per serving, he won't count out scallions to feed a thousand folks, he'll calculate the weight of 2,000 scallions and put baskets of them on the scale until he gets that weight.
- Chemists are in the same boat.
  - Our recipe calls for 2 atoms of hydrogen and 1 of oxygen per serving, to make water. But we need 10^{23} servings to fill a thimble with water.
  - Just like a banker needs to know the weights of quarters and pennies, we need to know the weights of carbon atoms, nitrogen atoms, and hydrogen atoms. We need the weights of our elements.

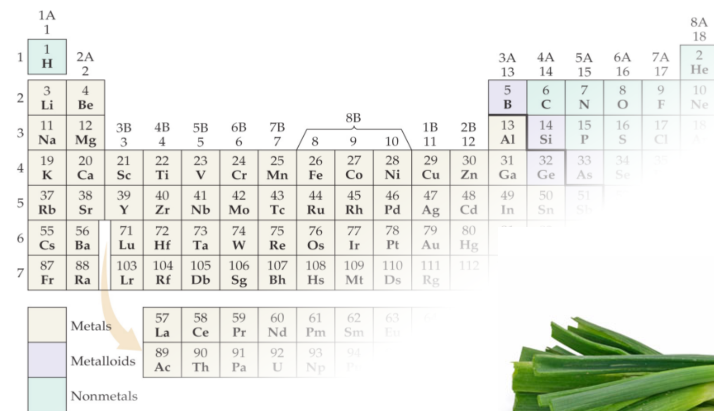


Counting by Weight

- Every flavor atom is made of neutrons & protons.
  - It's convenient when we're working on a molecular scale to have a unit of weight about the size of a neutron or proton.
  - We call that unit amu (atomic mass unit).
  - Most interesting molecules are made of carbon.
  - The most common isotope of carbon is made almost entirely of 6 protons and 6 neutrons.
  - An amu is defined as:

    exactly $\frac{1}{12}$ the mass of Carbon-12

  - 1 amu is measured to be 1.6606×10^{-24} g.
  - ➡ (you don't need to memorize this)
- A chef weighing tomatoes doesn't use the weight of the largest tomato or the smallest. He uses the average weight of a tomato.
- Not all carbon atoms weigh the same, if we're weighing out carbon atoms we want to use average weight of a carbon atom.
- How do we get the average weight?



Weighted Averages

- How do you find the average mass of a tomato?
- If you have two tomatoes, you add their mass and divide by the number of tomatoes.

200 grams

100 grams

$$\frac{200g+100g}{2}=150g$$

- If you have a lot of tomatoes, it might be easier to multiply the amount of tomatoes you have of each mass by that value rather than add them one at a time.
- The number of tomatoes at each mass over the total number of tomatoes is also the percent at each mass – if 8 of your 10 tomatoes is 100 grams, that's 80% of your tomatoes.

$$\frac{2x200g+8x100g}{10}$$

$$=\frac{2}{10}x200g+\frac{8}{10}x100g$$

$$=20\% \text{ of } 200g+80\% \text{ of } 100g$$

$$= 0.20 \times 200g + 0.80 \times 100g$$

$$=40g+80g$$

$$=120g$$

- If you have so many tomatoes you don't know the total number, you can take a sample of them and determine the percent that are 100 g and 200 g in your sample.
- As long as the sample is a good representation of the total, it produces the same average mass as if we added the mass of all the tomatoes and divided by the total.
- We weight the heavier value 80% because those tomatoes occur four times as often as the tomatoes we apply the 20% weighting factor to.
- We might not know how many tomatoes we have, but if we know 20% of them mass 200 g and 80% mass 100 g we know that if we pick up a random bucket of tomatoes the average mass for that bucket will be 120g each.

$$20\% \text{ of } 200g + 80\% \text{ of } 100g$$

$$= 0.20 \times 200g + 0.80 \times 100g$$

$$= 40g + 80g$$

$$= 120g$$

Average Atomic Mass

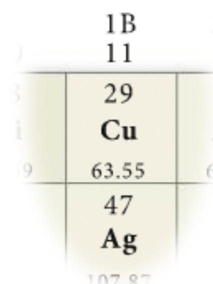
- The periodic table only reports one mass for each element, how does that work if each element has isotopes of different masses?
- The ratio of naturally occurring isotopes of each element is known.
- Every time we pour out a sample of copper, we know 69% of it's atoms are copper-63 and 31% are copper-65.
- Everytime.
- So we don't care what the mass of each isotope is, just what the mass – on average – of a copper atom.
- The periodic table represents an average atomic mass for that element.

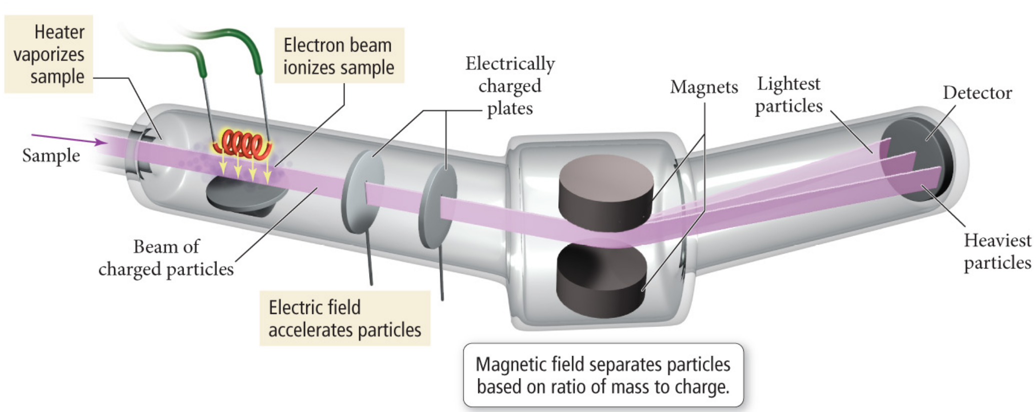
Isotope	Isotopic mass (amu)	Abundance (%)	Average atomic mass (amu)
$^{63}_{29}Cu$	62.9298	69.09	63.55
$^{65}_{29}Cu$	64.9278	30.91	

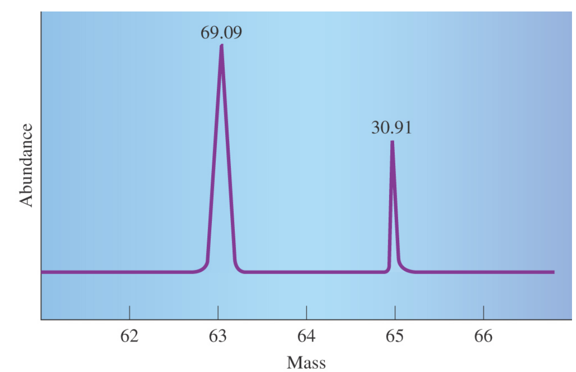
62.9298 amu × 0.6909 = 43.48 amu

64.9278 amu × 0.3091 = 20.07 amu

63.55 amu







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- Illustrative Examples
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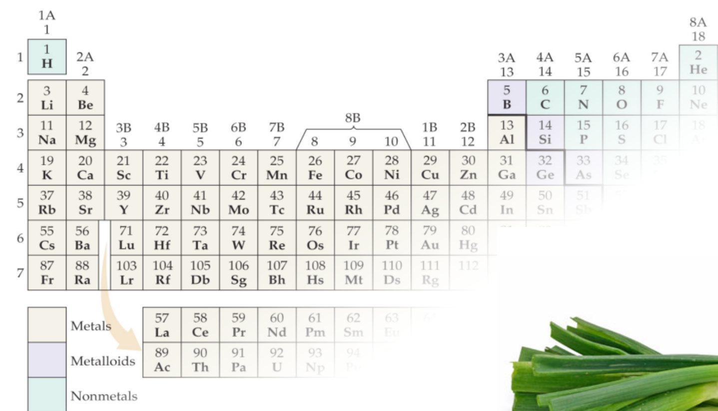
        - more...



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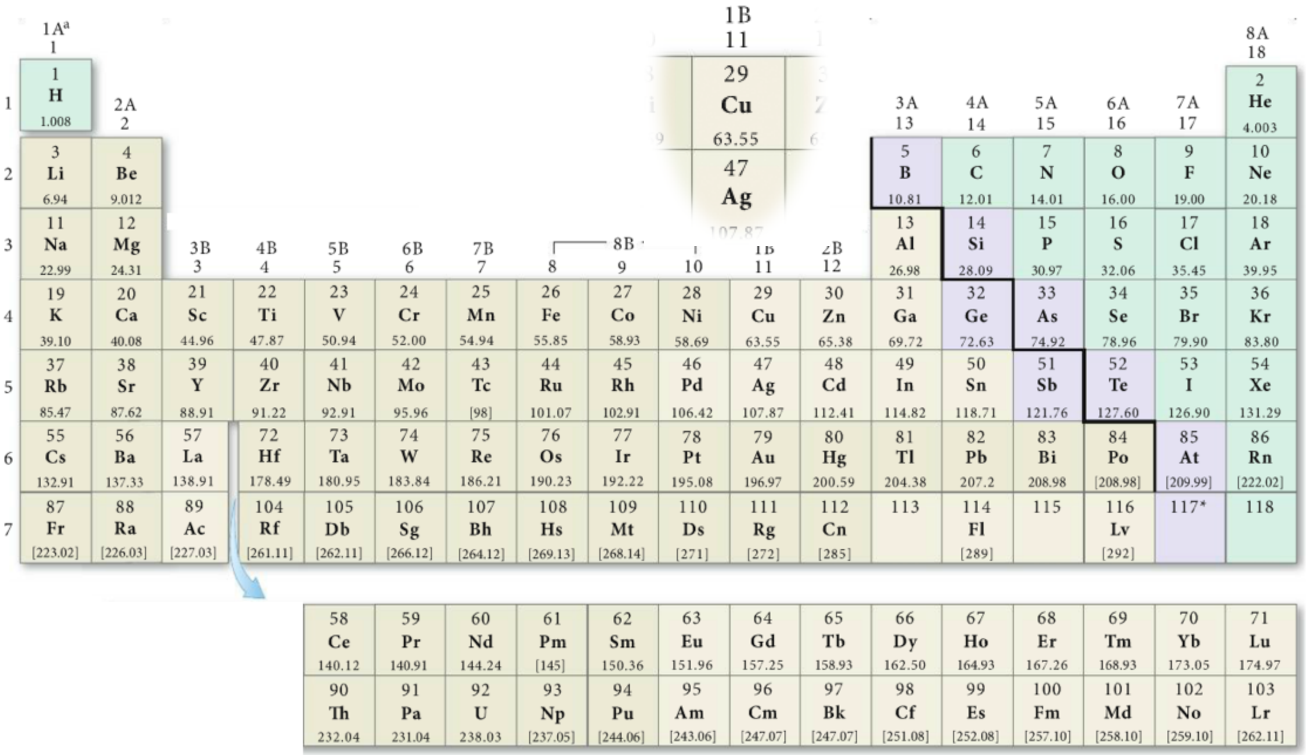
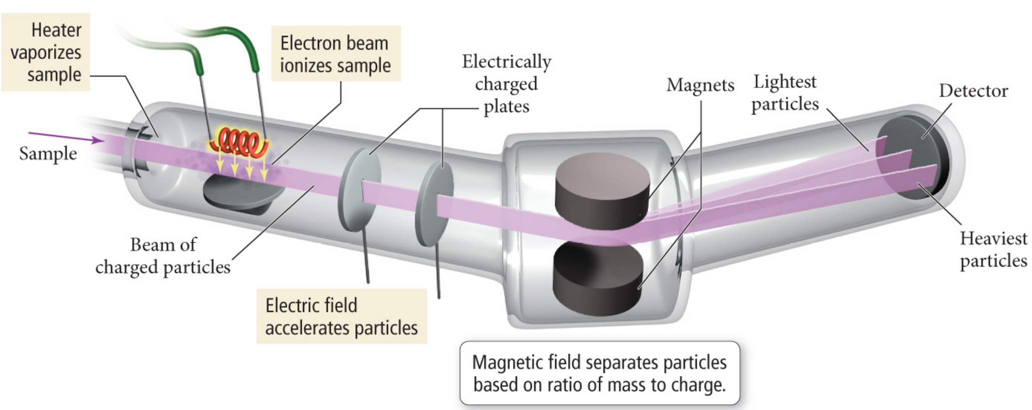
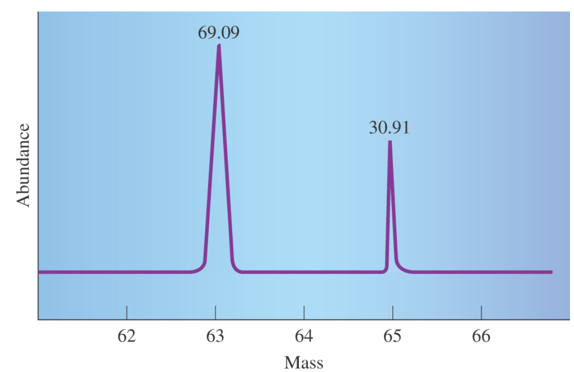




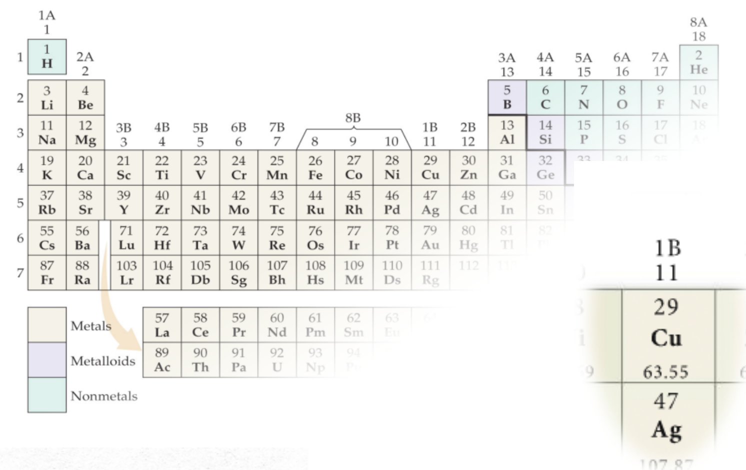


Average Atomic Mass

1 Cu = 63.55 amu

What's the weight of one copper atom?

$$1\ Cu \times \frac{63.55\ amu}{1\ Cu} = 63.55\ amu$$



What's the weight of 17 copper atoms?

$$17\ Cu \cdot \frac{63.55\ amu}{1\ Cu} = 1,080.35\ amu$$

(counting # ∞ s.f.) 4 s.f. = 1,080 amu

1080. amu or 1.080 x 10^3 amu

How many copper atoms in two pennies?

(a penny weighs about 3.0 grams, an amu = 1.6606 x 10^{-24} g)

penny → g → amu → atoms

$$2\ penny \cdot \frac{3.0\ g}{1\ penny} \cdot \frac{1\ amu}{1.6606 \times 10^{-24}\ g} \cdot \frac{1\ Cu}{63.55\ amu} = 5.68553 \times 10^{22}$$

(counting # ∞ s.f.) 2 s.f. 5 s.f. 4 s.f.

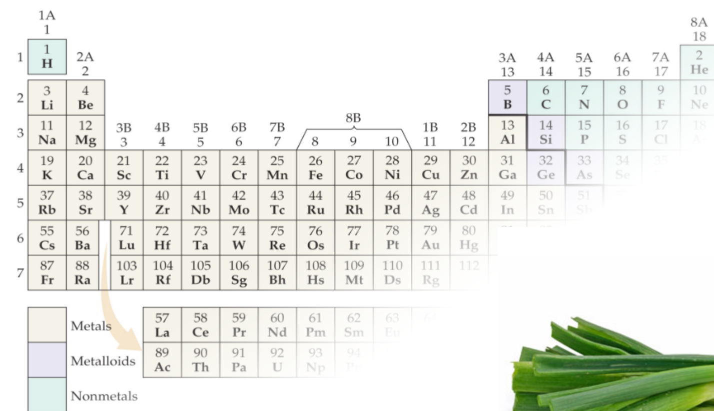
5.7 x 10^{22} Cu copper atoms

Problems:
- **we need a ratio of atoms for our recipes (ie H_2O)**
- **in the lab we want to use grams**
- **we don't want to have to convert to amu every time we need to count atoms**
- **and x10^{24} is awkward number to work with anyway.**

The Chemist's Dozen

- A recipe doesn't always list ingredients by single servings. Sometimes it uses dozens, score, or gross.
- When you're cooking for large groups, your recipe might call for 4 dozen eggs or 6 gross of dumplings.
    - 1 dozen = 12 singles
    - 1 score = 20 singles
    - 1 gross = 144 singles
- Working with dozens instead of singles let's a chef prepare on a scale 12x his design scale.
- We need a chemists dozen.
- We need to go from amu things (1 amu = 1.6606×10^{-24} g) to gram things (lab scale).
    - 1 gram ÷ 1 amu (in grams) = 6.022×10^{23}
    - **1 gram ÷ 1.661×10^{-24} grams = 6.022×10^{23}**
- We call 6.022×10^{23} singles a mole.
- It's the chemists dozen. We abbreviate mole as mol.
- A mol is a measurement, we will determine it to 4 sig figs and use it with 4 sig figs for most of this class.
- The number of singles in a mol is called Avogadro's Number.
- A mol is officially defined as the number of Carbon-12 atoms in 12 grams of pure Carbon-12 (you get the same number)

the tool for going between molecular scale (amu) and lab scale (grams).

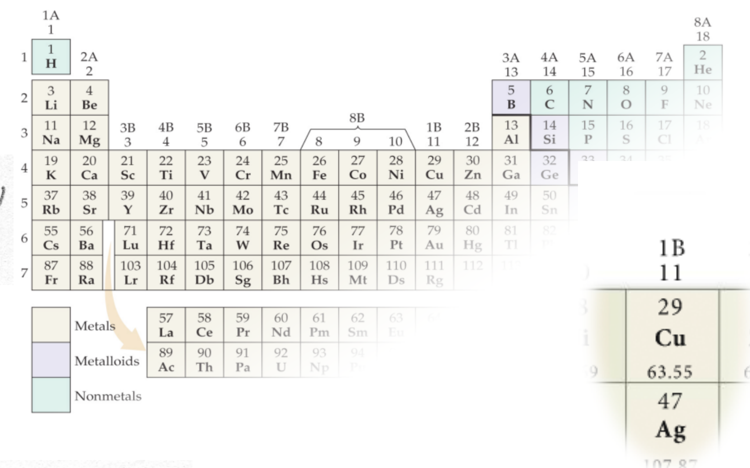


The Chemist's Dozen

1 mol = 6.022×10^{23} singles

How many atoms in exactly 1 mol Copper (Cu)?

exactly 1 mol Cu $\cdot \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 6.022 \times 10^{23}$ atoms Cu



How many atoms in 2.53 mol Copper (Cu)?

2.53 mol Cu $\cdot \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 1.52357 \times 10^{24}$ atoms Cu

1.52×10^{24} atoms Cu

How many mol Cu in 30.5 grams Cu?

g → amu → atoms → mol   There's an easier way!

How many Cu atoms in 30.5 grams Cu?

Atomic Weights / Molar Weights

- Weights are listed in the periodic table without units.
- The weight listed is the average mass of one atom of each element, in amu.
-

  1 gram ÷ 1.6606 x 10^{-24} grams = 6.022 x 10^{23}

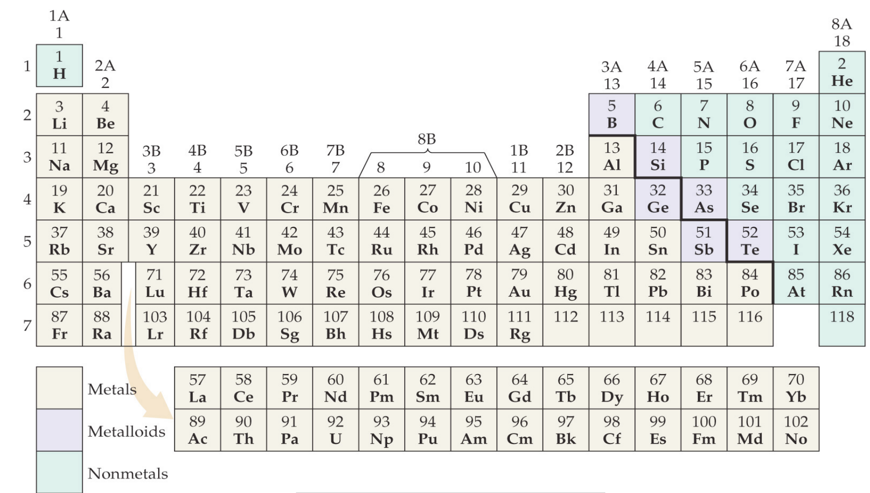
  1 gram ÷ 1 amu = 1 mol

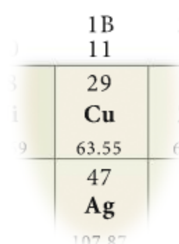
  1 gram = 1 mol x 1 amu

- That means:

  1 mol of *anything* will weigh in grams,
  what a single of that *anything* weighs in amu.

- If a cat weighs X amu, a mol of cats weighs X grams.
- That means each weight in the periodic table is:
  - the weight of 1 atom of that element, in amu
  - the weight of 1 mol of that element, in grams
- Reading from the periodic table...
  - a hydrogen atom (H) weighs 1.008 amu
  - a mol of hydrogen atoms (H) weigh 1.008 g
  - a copper atom (Cu) weighs 63.55 amu
  - a mol of copper atoms (Cu) weighs 63.55 g





1 H = 1.008 amu

1 mol H = 1.008 g

1 Cu = 63.55 amu

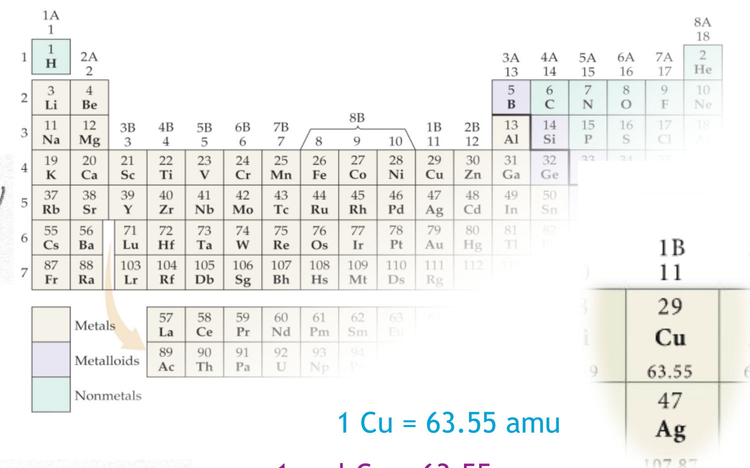
1 mol Cu = 63.55 g

Counting by Weight

1 mol = 6.022×10^{23} singles

How many atoms in exactly 1 mol Copper (Cu)?

$$\text{exactly } 1 \text{ mol Cu} \cdot \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = \boxed{6.022 \times 10^{23} \text{ atoms Cu}}$$



1 Cu = 63.55 amu

1 mol Cu = 63.55 g

How many atoms in 2.53 mol Copper (Cu)?

$$2.53 \text{ mol Cu} \cdot \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 1.52357 \times 10^{24} \text{ atoms Cu}$$

$$\boxed{1.52 \times 10^{24} \text{ atoms Cu}}$$

How many mol Cu in 30.5 grams Cu?

g → mol

$$30.5 \text{ g Cu} \cdot \frac{1 \text{ mol}}{63.55 \text{ g}} = 4.79937 \times 10^{-1} \text{ mol Cu}$$

3 sf, 4 sf

$$\boxed{0.480 \text{ mol Cu}}$$

How many Cu atoms in 30.5 grams Cu?

g → mol → atoms

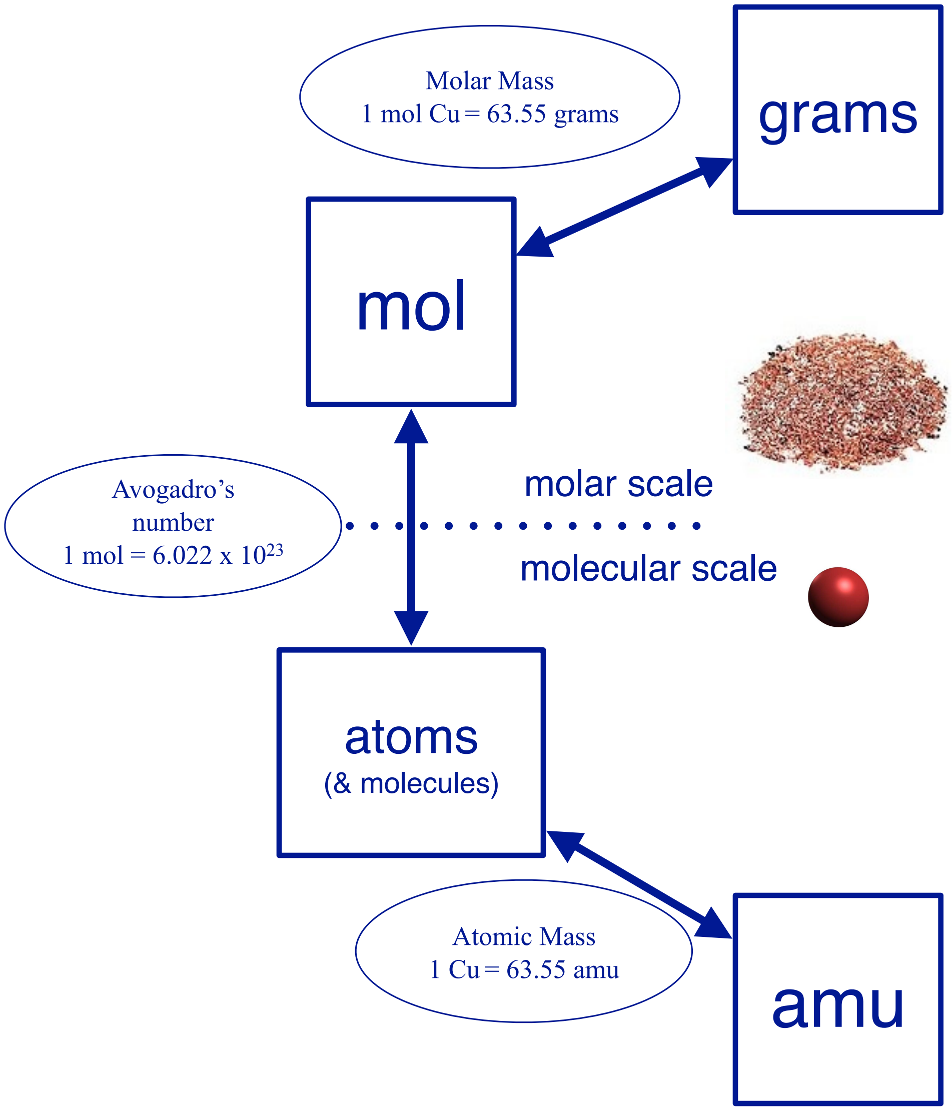
$$30.5 \text{ g Cu} \cdot \frac{1 \text{ mol}}{63.55 \text{ g}} \cdot \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 2.8901809 \times 10^{23} \text{ atoms}$$

3 sf, 4 sf, 4 sf

$$\boxed{2.89 \times 10^{23} \text{ atoms Cu}}$$

Mapping it Out

- Let's map it out.
- Places we go between:
  - molecular scale: atoms, amu
  - molar scale: mol, grams (and more are coming…)
- What gets us there (conversion factors)
  - Avogadro's Number
  - Molar Weight (aka Molar Mass)
  - Atomic Weight (aka Atomic Mass)
- Some Possible Conversions
  - How do we go from grams to atoms?
    - g → mol → atoms
      - molar mass; Avogadro's number
  - How do we go from atoms to mol?
    - atoms → mol
      - Avogadro's Number
  - How do we go from atoms to grams?
    - atoms → mol → grams
      - Avogadro's Number; molar mass
  - How do we go from grams to atoms?
    - mol → grams
      - molar mass



Questions?