

Lecture 1 - 5th September 2024

density: $\rho = \frac{m}{V}$ → water density: 997 kg/m^3

Significant Digits

- all nonzero digits are significant!
- for quantities less than one, any zero preceding the first non-zero digit is not significant.
↳ eg: 0.00237 has 3 sig figs.
- in quantities greater than one, any zero following the last non-zero digit is not significant, unless there is a decimal point.
↳ eg: 200 has only one sig fig, while 200.00 has five.
- when adding or subtracting, retain digits to the magnitude of the value with the least significant digits.
- when multiplying or dividing, keep the same number of digits as the value with the least significant digits.

Stoichiometry:

Example: what is $7.2 \frac{\text{Btu}}{\text{lbm}^\circ\text{F}}$ expressed in units of $\text{Jg}^{-1}\text{K}^{-1}$?

↳ given: $1 \text{ J} = 9.4782 \cdot 10^{-4} \text{ Btu}$, $1 \text{ kg} = 2.20462 \text{ lbm}$, $1 \text{ K} = 1.8^\circ\text{F}$

$$\rightarrow 7.2 \frac{\text{Btu}}{\text{lbm}^\circ\text{F}} \cdot \frac{1 \text{ J}}{9.4782 \cdot 10^{-4} \text{ Btu}} \cdot \frac{2.20462 \text{ lbm}}{1 \text{ kg}} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} \cdot \frac{1.8^\circ\text{F}}{1 \text{ K}}$$

$$= 30.1 \frac{\text{J}}{\text{gK}}$$

Example: you are interested in calculating the annual reduction in CO_2 emissions by replacing your car with an all-electric vehicle. Assume that you drive an annual distance 14000 km. Your car has a fuel efficiency 28 miles per gallon. The electric car has an average energy consumption of 157 Wh/km.

↳ Data for CO_2 emission: 25g CO_2 /kWh, 2.3 kg CO_2 /L

$$\begin{aligned} \text{Gas: } 14000 \frac{\text{km}}{\text{year}} &\cdot \frac{6.21371 \cdot 10^{-4} \text{ mi}}{1 \text{ m}} \cdot \frac{1000 \text{ m}}{1 \text{ Km}} \cdot \frac{1 \text{ gal}}{28 \text{ mi}} \\ &\cdot \frac{1000 \text{ L}}{264.172 \text{ gal}} \cdot \frac{2.3 \text{ kg CO}_2}{1 \text{ L}} = 2705 \frac{\text{kg CO}_2}{\text{year}} \end{aligned}$$

$$\begin{aligned} \text{Electric: } 14000 \frac{\text{km}}{\text{year}} &\cdot \frac{157 \text{ Wh}}{1 \text{ Km}} \cdot \frac{25 \text{ g CO}_2}{1 \text{ kWh}} \cdot \frac{1 \text{ kWh}}{1000 \text{ Wh}} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} \\ &= 55 \frac{\text{kg CO}_2}{\text{year}} \end{aligned}$$

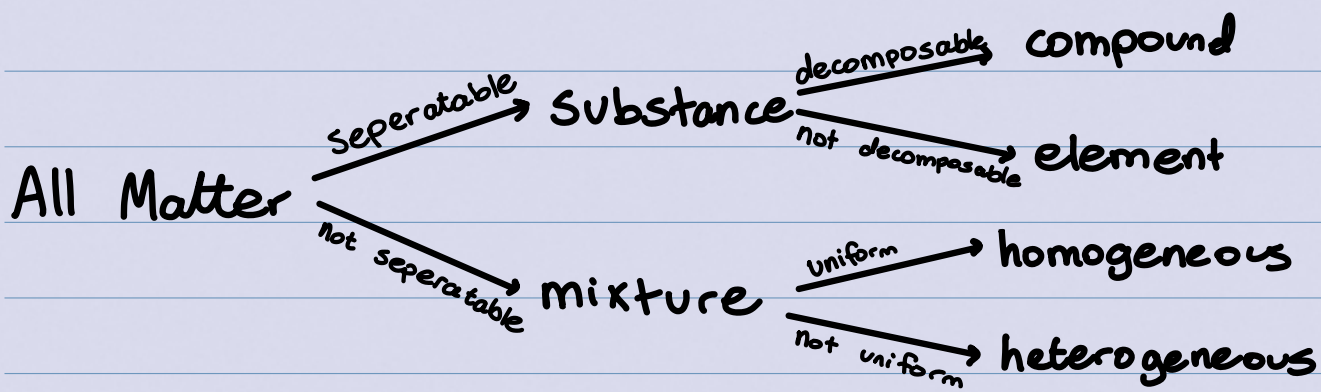
$$\text{reductions: gas - electric} = 2705 - 55 = 2650 \frac{\text{kg CO}_2}{\text{year}}$$

Temperature Conversions:

- $T(^{\circ}\text{F}) = 1.8 T(^{\circ}\text{C}) + 32$
- $T(^{\circ}\text{C}) = \frac{T(^{\circ}\text{F}) - 32}{1.8}$
- $T(\text{K}) = T(^{\circ}\text{C}) + 273.15$
- $T(^{\circ}\text{R}) = 1.8 T(\text{K})$
- $T(^{\circ}\text{R}) = T(^{\circ}\text{F}) + 491.67$

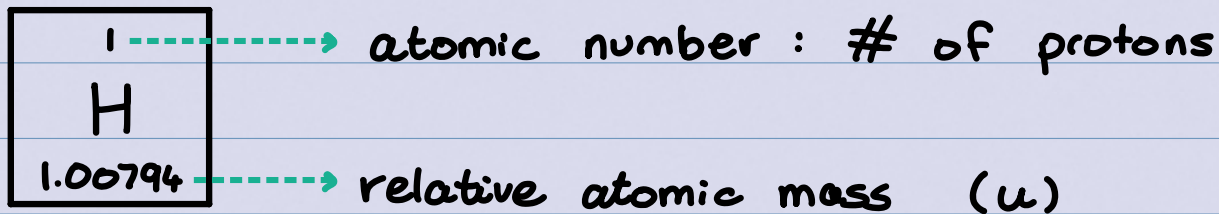
all given on exams!

Classification of Matter



Example: classify the following:

- a) a piece of metal: can't determine
- b) distilled water: pure substance ↗ bc when seperated, both are still H₂O!
- c) distilled water w/ ice: pure substance
- d) oil on water: heterogeneous
- e) coffe w/ cream and sugar: heterogenous



↳ # protons + # electrons = relative atomic mass!

Example: the mass number and the atomic number of an atom that contains 15 electrons, 15 protons, and 16 neutrons are, respectively:

15 protons means that the atomic number is 15.
the mass is # protons + # neutrons = 15 + 16 = 31.

∴, 31 and 15.

- Atoms are most stable when their outer shell is completely empty or completely full.
↳ octet rule!
- Elements combine to form compounds by atomic bonding.
- Bonding involves electrons that are either shared (covalent bond) or transferred (ionic bond)
- Energy is absorbed or released when electrons are transferred or shared.

Example: copper has two major naturally occurring isotopes, which have 34 and 36 neutrons. Using the periodic table, determine the percentage abundance of copper atoms having 34 neutrons.

- From the periodic table, the atomic number of copper is 29 and the atomic mass is 63.546 amu.

$$\begin{aligned} \text{↳ mass of } {}^{63}\text{Cu} : 29 + 34 &= 63\text{u} \\ \text{mass of } {}^{65}\text{Cu} : 29 + 36 &= 65\text{u} . \end{aligned}$$

$$63.546\text{u} = x \cdot 63\text{u} + (1-x) \cdot 65\text{u}$$

$$\therefore, x = 0.727 = 72.7\%$$

• **Electronegativity**: a measure of how strongly an atom competes for electrons in bonds formed with other atoms.

↳ increases rightwards in the periodic table!

Ionic Compounds: one atom donates one or more electron to another atom.

- Usually a metal combined with a non-metal.

- Dissociate into ions when dissolved in water.

↳ Eg: NaCl, ordered crystal of Na^+ and Cl^- ions.

Covalent Compounds: one or more atoms share one or more electrons.

- Large, organic molecules are usually covalent.

- Do not dissociate into ions when dissolved in water.

↳ Eg: CH_4 , $\text{C}_6\text{H}_{12}\text{O}_6$

- Two or more atoms, joined by covalent bonds, may form a charged polyatomic ion.

↳ Eg: ammonium (NH_4^+), carbonate (CO_3^{2-}), etc.

- **Molecular formula**: shows the actual number of atoms in a molecule.

- **Empirical formula**: gives the relative number of atoms in a molecule.

- The molecular formula can be obtained from the

empirical formula if the molecular mass of the compound is known.

Example: C_6H_{12} is the molecular formula, while CH_2 is the empirical formula of hexene.

Example: A compound has the empirical formula CH_2O and a molecular weight of 150.13 u . What is the molecular formula?

↳ weight of C : 12.011

weight of H_2 : $1.008 \cdot 2 = 2.016$

weight of O : 15.999

$$\therefore, \text{weight of } CH_2O = 12.011 + 2.016 + 15.999 \\ = 30.026 \text{ u}.$$

We know that the compound should have a molecular mass of 150.13 u , so how many times does one CH_2O go into it?

$$150.13 \div 30.026 = 5.$$

\therefore , compound is $5CH_2O = C_5H_{10}O_5$.