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 $7.2\ \frac{\text{Btu}}{\text{lbm}\,°\text{F}}$ expressed in units of $\text{J}\,\text{g}^{-1}\,\text{K}^{-1}$?

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

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

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

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 effeciency 28 miles per gallon. The electric car has an average energy consumption of 157 Wh/km.

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

↳ Gas: $$14000 \frac{km}{year} \cdot \frac{6.21371 \cdot 10^{-4} mi}{1 m} \cdot \frac{1000 m}{1 km} \cdot \frac{1 gal}{28 mi} \cdot \frac{1000 L}{264.172 gal} \cdot \frac{2.3 kg CO_2}{1 L} = 2705 \frac{kg CO_2}{year}$$

Electric: $$14000 \frac{km}{year} \cdot \frac{157 Wh}{1 km} \cdot \frac{25 g CO_2}{1 kWh} \cdot \frac{1 kWh}{1000 Wh} \cdot \frac{1 kg}{1000 g} = 55 \frac{kg CO_2}{year}$$

reductions: gas - electric = $2705 - 55 = 2650 \frac{kg CO_2}{year}$

Temperature Conversions:

- $T(°F) = 1.8\, T(°C) + 32$
- $T(°C) = \frac{T(°F) - 32}{1.8}$
- $T(K) = T(°C) + 273.15$
- $T(°R) = 1.8\, T(K)$
- $T(°R) = T(°F) + 491.67$

all given on exams!

Classification of Matter

All Matter
- seperatable → Substance
  - decomposable → compound
  - not decomposable → element
- not seperatable → mixture
  - uniform → homogeneous
  - not uniform → heterogeneous

Example: classify the following:

a) a piece of metal: can't determine

b) distilled water: pure substance

c) distilled water w/ ice: pure substance ↳ bc when seperated, both are still H_2O!

d) oil on water: heterogeneous

e) coffe w/ cream and sugar: heterogenous

1
H
1.00794

1 → atomic number: # of protons

1.00794 → relative atomic mass (u)

↳ # 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 + # electrons = 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 occuring 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.

↳ mass of ^{63}Cu : $29 + 34 = 63u$
mass of ^{65}Cu : $29 + 36 = 65u$.

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

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

- **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, $C_6H_{12}O_6$

- 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.

- **Emperical 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 emperical 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

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

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

$150.13 \div 30.026 = 5.$

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