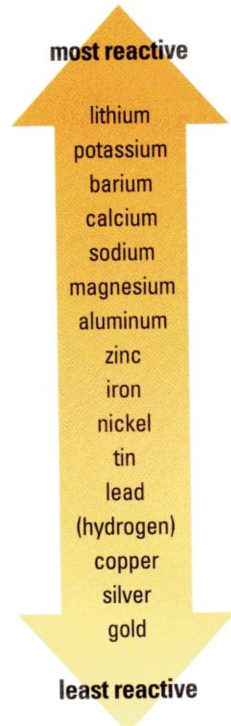
SCH3U FORMULA SHEET

$n = \frac{m}{M}$ $\qquad$ $N = nN_A$

Avogadros Number (N_A): 6.023×10^{23}



$C = \frac{n}{V}$ $\qquad$ $C_1V_1 = C_2V_2$

$ppm = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$

$ppb = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9$

$pH = -\log[H^+]$

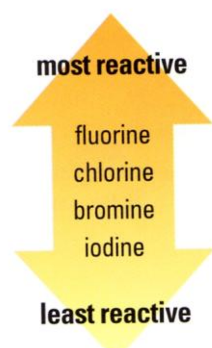
$[H^+] = 10^{-pH}$

STP = 273 K $\quad$ 101.325 kPa $\qquad$ R = 8.31 kPa•L/mol•K

STAP = 298 K $\quad$ 100 kPa

1 atm = 101.325 kPa = 760 torr = 760 mm Hg

$P_1V_1 = P_1V_2$ $\qquad$ $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$



$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $\qquad$ $PV = nRT$

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ $\qquad$ $P_{total} = P_1 + P_2$

Molar volume of a gas at STP (0^0C and 100.325 kPa) $\qquad$ V_{STP}= 22.4 L/mol

Molar volume of a gas at SATP (25^0C and 100.0 kPa) $\qquad$ V_{SATP} = 24.8 L/mol

SCH3U Cheat Sheet for SCH4U

When are Digits Significant?

- Non-zero digits are always significant. Thus, 22 has two significant digits, and 22.3 has three significant digits.
- With zeroes, the situation is more complicated:
  1. Zeroes placed before other digits are not significant; 0.046 has two significant digits.
  2. Zeroes placed between other digits are always significant; 4009 kg has four significant digits.
  3. Zeroes placed after other digits but behind a decimal point are significant; 7.90 has three significant digits.
  4. Zeroes at the end of a number are significant only if they are behind a decimal point as in (c). Otherwise, it is impossible to tell if they are significant. For example, in the number 8200, it is not clear if the zeroes are significant or not. The number of significant digits in 8200 is at least two, but could be three or four. To avoid uncertainty, use scientific notation to place significant zeroes behind a decimal point:

8.200×10^3 has 4 sig figs 8.20×10^3 has 3 sig figs 8.2×10^3 has 2 sig figs

Significant Digits in Multiplication, Division, Trig. functions, etc.

- In a calculation involving multiplication, division, trigonometric functions, etc., the number of significant digits in an answer should **equal the least number of significant digits** in any one of the numbers being multiplied, divided etc.
  - Multiplying 0.097 m^{-1} (two significant digits) and 4.73 m (three significant digits), the answer should have two significant digits.
- Keep One Extra Digit in Intermediate Answers
  - When doing multi-step calculations, *keep at least one more significant digit in intermediate results* than needed in your final answer.
- For instance, if a final answer requires two significant digits, then carry at least three significant digits in calculations. If you round-off all your intermediate answers to only two digits, you are discarding the information contained in the third digit, and as a result the *second* digit in your final answer might be incorrect. (This phenomenon is known as "round-off error.")

The Two Greatest Sins Regarding Significant Digits

- Writing more digits in an answer (intermediate or final) than justified by the number of digits in the data.
- Rounding-off, say, to two digits in an intermediate answer, and then writing three digits in the final answer.

Scientific Notation:

- Only one number before the decimal
- Positive exponents are large numbers (4.5×10^6 is 4500000)
- Negative exponents are small numbers (4.5×10^{-6} is 0.0000045)
- Entering on your scientific calculator:
  - Punch the number (the digit number) into your calculator.
  - **Push the EE or EXP button. Do NOT use the x (times) or 10^x button!!**
  - Enter the exponent number. Use the +/- button to change its sign.
  - Voila! Treat this number normally in all subsequent calculations.
  - To check yourself, multiply 6.0×10^5 times 4.0×10^3 on your calculator. Your answer should be 2.4×10^9.

Naming Rules (ionic)

- Write symbols in order (metals first then non metals)
- Write valences or charge above each symbol (multivalent metals look on Periodic Table)
- Crisscross charges to make them subscripts
- Reduce subscripts if needed (ie 2 2 becomes 1 1)
- Nonmetal ending is changed to –ide

  e.g. a compound composed of Fe^{2+} and Cl

  Fe^{2+} Cl^{1-} $FeCl_2$ iron (II) chloride

Naming Rules (polyatomic):

- Same rules as above, see table for list of names and charges

  e.g. K^{1+} ClO_3^- $KClO_3$ potassium chlorate

Naming Rules (hydrates):

- Use mono-, di- etc to indicate number of water molecules

  e.g. $CuSO_4 \bullet 5\ H_2O$ copper(II) sulphate **penta**hydrate

Naming Rules (molecular):

- Use mono-, di-, tri- etc to indicate number of atoms of each present in the compound.

  e.g. NO_2 nitrogen dioxide

Naming Rules (acids):

Name of Ion	Classic Acid Name	IUPAC Acid Name	Examples
___ ide	hydro ___ ic acid	aqueous hydrogen ___ide	HCl – **hydro**chlor**ic acid**
per___ate	per___ic acid	aqueous hydrogen per___ate	$HClO_4$ – **per**chlor**ic acid**
___ate	___ic acid	aqueous hydrogen ___ate	$HClO_3$ – chlor**ic acid**
___ite	___ous acid	aqueous hydrogen ___ite	$HClO_2$ – chlor**ous acid**
hypo___ite	hypo___ous acid	aqueous hydrogen hypo___ite	HClO – **hypo**chlor**ous acid**

Note: If you memorize the "ate" ions, you'll recognize a trend…add an oxygen you get the "per__ate" ion, take away one oxygen you get the "ite" ion, take away two oxygens you get the "hypo__ite" ion.

Reaction Types:

Reaction type	Reactants	Products
combustion	metal + oxygen	metal oxide
	nonmetal + oxygen	nonmetal oxide
	fossil fuel + oxygen	carbon dioxide + water
synthesis	element + element	compound
	element + compound	more complex compound
	compound + compound	more complex compound
decomposition	binary compound	element + element
	complex compound	simpler compound + simpler compound or simpler compound + element(s)
single displacement	A + BC	B + AC
double displacement	AB + CD	AD + CB

Mole and Concentration Formulas:

$n = \frac{m}{M}$ $C = \frac{n}{V}$

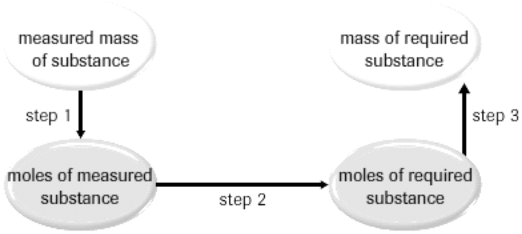
$$C_iV_i = C_fV_f$$

Stoichiometry:

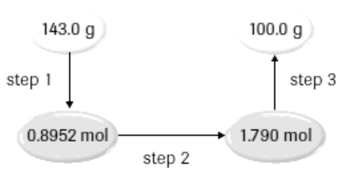
Calculating Mass of Reactants and Products

Begin with a balanced chemical equation, with the measured mass of reactant or product written beneath the correspon-ding formula.

1. Convert the measured mass into an amount in moles.
2. Use the mole ratio in the balanced equation to predict the amount in moles of desired substance.
3. Convert the predicted amount in moles into mass (See example, **Figure 4**).



$$Fe_2O_{3(s)} + 3\ CO_{(g)} \rightarrow 2\ Fe_{(s)} + 3\ CO_{2(g)}$$



Determining the Limiting Reactant

1. Write a balanced equation for the reaction.
2. Select one of the reactants and calculate the amount in moles *available*.
3. Use mole ratios in the balanced equation to calculate the amount in moles *needed* of the *other* reactants.
4. Calculate the *available* amount in moles of the *other* reactants. If the available amount of a reactant is more than sufficient, it is in excess. If the available amount is insufficient, it is limiting. (See example, **Figure 5**)

(a) $CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$