

CHE 102 Final Reference Sheet - Sean Yang

Temperature Conversion Factors

$$\frac{1K}{1.8^{\circ}R}, \frac{1K}{1.8^{\circ}F}, \frac{1^{\circ}C}{1.8^{\circ}F}, \frac{1^{\circ}C}{1.8^{\circ}R}, \frac{1^{\circ}C}{1K}, \frac{1^{\circ}F}{1^{\circ}R}$$

(Relative differences only)

Density

$$\rho = \frac{m}{v}$$

Mole Fraction and Percentage

$$x_A = \frac{n_A}{n_T} = \frac{n_A}{\sum_i n_i}$$

$$\% = x_A \times 100\%$$

Mass Fraction and Percentage

$$w_A = \frac{m_A}{m_T} = \frac{m_A}{\sum_i m_i}$$

$$\% = w_A \times 100\%$$

Average Molar Mass

$$M = \sum_i x_i M_i$$

Percentage Yield and Excess

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\%$$

$$\% \text{ excess} = \frac{\text{amount provided} - \text{amount required}}{\text{amount required}} \times 100\%$$

Boyle's Law (Constant n, T)

$$P \propto \frac{1}{v}$$

$$P_1 V_1 = P_2 V_2$$

Charles' Law (Constant n, P)

$$V \propto T$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Temperature in K

Avogadro's Law (Constant T, P)

$$V \propto n$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

At 0°C and 1 atm: 1 mol gas = 22.414 L

At 25°C and 1 atm: 1 mol gas = 24.465 L

Ideal Gas Law

$$PV = nRT$$

Temperature in K

Ideal Gas Assumptions:

- No intermolecular forces between molecules
- Gas molecules have no volume

Approximation valid at high T , low P

Gas Density

$$\rho = \frac{PM}{RT}$$

$$\text{at } 25^{\circ}\text{C and 1 atm: } \rho = \frac{M}{24.465 \text{ L} \cdot \text{mol}^{-1}}$$

Gas Concentration

$$C = \frac{P}{RT}$$

$$\text{at } 25^{\circ}\text{C: } C = \frac{P}{24.465 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1}}, \text{ where } P \text{ is in atm}$$

Dalton's Law of Partial Pressures

$$P = P_A + P_B$$

$$\text{Assuming } V = V_A = V_B$$

Dalton's Law of Partial Volumes

$$V = V_A + V_B$$

$$\text{Assuming } P = P_A = P_B$$

Dalton's Law (Mole Fraction)

$$y_A = \frac{n_A}{n} = \frac{P_A}{P} = \frac{V_A}{V}$$

$$n_A = y_A n$$

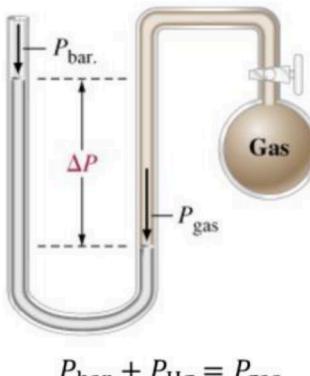
$$P_A = y_A P$$

$$V_A = y_A V$$

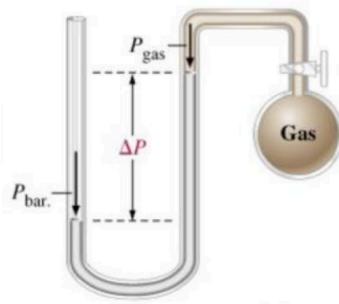
Measuring Pressure with an Open-End Manometer



(a) Gas pressure equal to barometric pressure



(b) Gas pressure greater than barometric pressure



(c) Gas pressure less than barometric pressure

$$\Delta P_{\text{mmHg}} = \Delta h_{\text{Hg}}$$

Kinetic Molecular Theory Postulates

- Gas particles are in constant, random motion
 - Collisions between gas particles and container walls are elastic
 - The average kinetic energy of gas particles is proportional to temperature in K
 - Gas particles exert no forces on each other (no attraction or repulsion)

Reaction Rate Constant Units

Note: Q_p and K_p should be calculated using bar units for pressure (1 bar = 10^5 Pa = 750 mmHg)

Order	Units of k
0	$\text{mol} \cdot L^{-1} \cdot s^{-1}$
1	s^{-1}
2	$L \cdot \text{mol}^{-1} \cdot s^{-1}$
n	$L^{n-1} \cdot \text{mol}^{1-n} \cdot s^{-1}$

Kinetic Plots

Order	Plot of	Slope
0	$[A]$ vs. t	$-k$
1	$\ln[A]$ vs. t	$-k$
2	$\frac{1}{[A]}$ vs. t	k

Integrated Rate Laws

Note: Pressure can be used in place of concentration for gases since concentration is proportional to pressure.

Order	Rate Law	Integrated Rate Law	Half-Life
0	rate = k	$[A] = [A]_0 - kt$	$t_{\frac{1}{2}} = \frac{[A]_0}{2k}$
1	rate = $k[A]$	$\ln[A] = \ln[A]_0 - kt$	$t_{\frac{1}{2}} = \frac{\ln 2}{k}$
2	rate = $k[A]^2$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$	$t_{\frac{1}{2}} = \frac{1}{k[A]_0}$

Henry's Law and Raoult's Law Units

C_A = concentration of gas in solution ($\frac{\text{mol}}{\text{L}}$)

k_A = Henry's law constant ($\text{mol} \cdot L^{-1} \cdot \text{atm}^{-1}$)

H_A = Henry's law constant (atm)

P_A = partial pressure of gas above solution (atm)

x_A = mole fraction of solvent in solution

Nernst Equation at 25°C

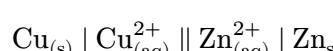
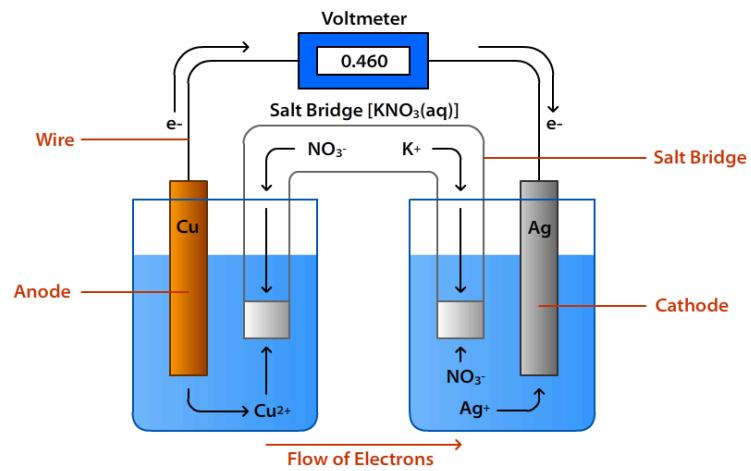
$$E = E^o - \left(\frac{0.0592}{n} \right) \log Q$$

Where n = number of moles of electrons transferred

E^o = standard cell potential (V)

Q = reaction quotient

Galvanic cell



Anode (oxidation): left side

Cathode (reduction): right side