Unit 6: SOLUTIONS AND ACIDS & BASES

Chapter 18: Solutions

18.1: Properties of Solution

Solubility: - the amount of solute that can dissolve in a given amount of solvent at a specific temperature.
- usually measures in g / 100 g of solvent.

1. **Unsaturated Solution**: - when a solvent can still dissolve more solute at a particle temperature.
2. **Saturated Solution**: - when a solvent cannot dissolve anymore solute.
   - the rate of solvation is the same as the rate of crystallization.
3. **Supersaturated Solution**: - when a saturated solution is heated and more solute is dissolved, which is then followed by a slow cooling without any recrystallization.
   - the density and the viscosity of a supersaturated solution is higher than a saturated solution. (Example: Sugar Syrup – Supersaturated Sugar Water)
4. **Miscible**: - when two liquids can dissolve each other.
   (Example: Vinegar-Water Solution – Both have polar bonds.)
5. **Partially Miscible**: - when two liquids are slightly soluble in each other.
   (Example: Anaesthetic Diethyl ether-Water Solution – weak polar versus strong polar bonds.)
6. **Immiscible**: - when two liquids cannot dissolve in each other. They form a mechanical (heterogeneous) mixture.
   (**Example**: Italian Salad Dressing – Vinegar and Oil: polar versus non-polar bonds.)

Factors Affecting Solubility

1. **Temperature**: - as temperature increases, solubility of solid solutes generally increases.
   - as temperature increases, solubility of gaseous solutes decreases in aqueous solution.

   Temperature ↑ **Solid Solute Solubility ↑** **Gas Solute Solubility ↓**

2. **Partial Pressure**: - as pressure increases, solubility of gas solutes generally increases. (More force is pushed down to force gas particles to dissolve in a denser liquid solvent).

   Pressure ↑ **Gas Solute Solubility ↑**

3. **Agitation (Stirring)**: - as agitation increases, solubility of solid solutes generally increases.

   Agitation ↑ **Solid Solute Solubility ↑**

4. **Particle Size**: - as the particle size of the solute increases, the rate of solubility decreases.

   Particle Size ↑ **Rate of Solubility ↓**

Example 1: A maximum 6.70 g of carbon dioxide gas is dissolved in 2000 g of water at 0°C under normal atmosphere pressure (101.3 kPa). Calculate the solubility of carbon dioxide gas in g / 100 g of H_2O.

$$\textbf{Solubility} = \frac{\textbf{Mass of Solute}}{\textbf{100 g of Solvent}}$$

$$\text{Solubility} = \frac{6.70\text{ g }CO_2}{2000\text{ g water}} \times \frac{x\text{ g }CO_2}{100\text{ g water}}$$ (cross-multiply)

$$2000x = (6.70)(100)$$

$$x = \frac{(6.70)(100)}{2000} = \mathbf{0.335\text{ g }CO_2}$$

$$\boxed{\textbf{Solubility} = \mathbf{0.335\text{ g }CO_2 / 100\text{ g of water}}}$$

Example 2: The solubility of sucrose (cane sugar) is 230.9 g / 100 g of H_2O. Determine the amount of solvent needed to dissolve 3.00 kg of cane sugar.

$$\text{Solubility} = \frac{230.9\text{ g sucrose}}{100\text{ g }H_2O} = \frac{3000\text{ g sucrose}}{x\text{ g }H_2O} \leftarrow \mathbf{3.00\text{ kg} = 3000\text{ g}}$$

(cross-multiply)

$$230.9x = (100)(3000)$$

$$x = \frac{(100)(3000)}{230.9} = 1299.263751\text{ g water}$$

$$\boxed{\mathbf{x = 1.30\text{ kg of }H_2O}}$$

Henry's Law: - states that the solubility of gas in a solution is directly proportional to the pressure above the solution.

$$\mathbf{S \propto P}$$

$$\mathbf{S = kP} \quad \text{or} \quad \mathbf{\frac{S}{P} = k}$$

***k* = constant of proportion**

Pressure ↑ Solubility of Gas ↑

Henry's Law

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S_1 = Solubility at Initial Condition
P_1 = Pressure of Initial Condition
S_2 = Solubility at Final Condition
P_2 = Pressure of Final Condition

Example 3: Some health food vendors claim that drinking oxygenated beverage will increase physical peformance. Suppose an oxygenated drink has a oxygen solubility of 3.17 g/L is under 7.50 atm of pressure. If the partial pressure of oxygen in the atmosphere 0.20947 atm, what will be the solubility of oxygen in the beverage if it was left open for a long time?

$P_1 = 7.50$ atm
$S_1 = 3.17$ g/L
$P_2 = 0.20947$
$S_2 = ?$

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

$$\frac{S_1P_2}{P_1} = S_2$$

As Pressure ↓, Solubility ↓

$$S_2 = \frac{(3.17\text{ g/L})(0.20947\ \cancel{\text{atm}})}{(7.50\ \cancel{\text{atm}})}$$

$$\boxed{\mathbf{P_2 = 0.0885\text{ g/L}}}$$

18.2: Concentration of Solution

Molarity (M): - moles of solute per volume of solution in Litres (mol/L).
- commonly referred to as **molar concentration (C)**.

Molar Concentration:

$$\text{Molarity (mol/L)} = \frac{\text{moles (mol)}}{\text{Volume (L)}} \qquad C = \frac{n}{V} \qquad 1\ \text{mol/L} = 1\ \text{M}$$

C = Molar Concentration **n = moles** **V = Volume**

Example 1: 3.46 g of copper (II) nitrate is dissolved in 250.0 mL of water. Calculate the molarity of the solution formed.

$m = 3.46\ \text{g}\ Cu(NO_3)_2$

$$n = \frac{3.46\ \cancel{\text{g}}}{187.564\ \cancel{\text{g}}/\text{mol}}$$

$n = 0.0184470367\ \text{mol}\ Cu(NO_3)_2$

$V = 250.0\ \text{mL} = 0.2500\ \text{L}$

$\mathbf{C = ?}$

$$C = \frac{n}{V} = \frac{0.0184470367\ \text{mol}}{0.2500\ \text{L}}$$

$$\boxed{\mathbf{C = 0.0738\ mol/L\ or\ 73.8\ mmol/L}}$$

Example 2: Determine the mass of sodium dichromate needed for 500.0 mL of 0.0300 M.

$V = 500.0\ \text{mL} = 0.5000\ \text{L}$

$C = 0.0300\ \text{mol/L}$

$M = 261.98\ \text{g/mol}\ Na_2Cr_2O_7$

$\mathbf{n = ?}$ $\mathbf{m = ?}$

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

$n = CV = (0.0300\ \text{mol}/\cancel{\text{L}})(0.5000\ \cancel{\text{L}})$

$\mathbf{n = 0.0150\ mol}$

$m = nM = (0.0150\ \cancel{\text{mol}})(261.98\ \text{g}/\cancel{\text{mol}})$

$$\boxed{\mathbf{m = 3.93\ g}}$$

Example 3: The Palmense Mineral Water from the city of Fermo in Italy has a sodium hydrogen carbonate concentration of 5.207 mmol/L. What volume of this mineral water will contain 4.00 mg of sodium hydrogen carbonate?

$m = 4.00\ \text{mg}\ NaHCO_3$

$$n = \frac{4.00\ \cancel{\text{mg}}}{84.0079\ \cancel{\text{g}}/\text{mol}}$$

$n = 0.0476145696\ \text{mmol}\ NaHCO_3$

$C = 5.207\ \text{mmol/L}$

$\mathbf{V = ?}$

$$C = \frac{n}{V} \qquad V = \frac{n}{C} = \frac{0.0476145696\ \cancel{\text{mmol}}}{5.207\ \cancel{\text{mmol}}/\text{L}}$$

$$\boxed{\mathbf{V = 0.00914\ L\ or\ 9.14\ mL}}$$

Dissociation: - when ionic compounds completely dissolve in water (100% soluble), the ionic bonds are severed and the ions "swim" freely in the new aqueous environment.

Chemical Dissociation Equation

$$A_xB_{y\,(s)} \rightarrow xA^{y+}_{(aq)} + yB^{x-}_{(aq)}$$

$[A^{y+}]$ = Molar Concentration of Cation A^{y+} **$[B^{x-}]$ = Molar Concentration of Anion B^{x-}**

Example 4: Write the chemical dissociation equation for the following ionic compounds when they dissolve in water.

a. $NaCl_{(s)}$

$$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

b. $Cu(NO_3)_{2\,(s)}$

$$Cu(NO_3)_{2\,(s)} \rightarrow Cu^{2+}_{(aq)} + 2\ NO_3^-_{(aq)}$$

Example 5: Calculate the molar concentration for each ion when the following ionic compounds dissolve in water.

a. 0.150 M of $NaOH_{(aq)}$

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$[NaOH] = 0.150\ M$

$$[Na^+] = 0.150\ \cancel{mol}/L\ \cancel{NaOH_{(s)}} \times \frac{1\ mol\ Na^+}{1\ \cancel{mol}\ \cancel{NaOH_{(s)}}} \qquad \boxed{[Na^+] = 0.150\ mol/L}$$

$$[OH^-] = 0.150\ \cancel{mol}/L\ \cancel{NaOH_{(s)}} \times \frac{1\ mol\ OH^-}{1\ \cancel{mol}\ \cancel{NaOH_{(s)}}} \qquad \boxed{[OH^-] = 0.150\ mol/L}$$

b. 3.45 g of potassium phosphate in 500 mL of water.

$$K_3PO_{4\,(s)} \rightarrow 3\ K^+_{(aq)} + PO_4^{3-}_{(aq)}$$

$m = 3.45\ g\ K_3PO_4$

$$n = \frac{3.45\ \cancel{g}}{212.274\ \cancel{g}/mol}$$

$n = 0.0162525792\ mol\ K_3PO_4$

$V = 500\ mL = 0.500\ L$

$C = ?$

$$C = \frac{n}{V} = \frac{0.0162525792\ mol}{0.500\ L}$$

$[K_3PO_4] = 0.0325\ mol/L$

$$[K^+] = 0.0325\ \cancel{mol}/L\ \cancel{K_3PO_{4\,(s)}} \times \frac{3\ mol\ K^+}{1\ \cancel{mol}\ \cancel{K_3PO_{4\,(s)}}} \qquad \boxed{[K^+] = 0.0975\ mol/L}$$

$$[PO_4^{3-}] = 0.0325\ \cancel{mol}/L\ \cancel{K_3PO_{4\,(s)}} \times \frac{1\ mol\ PO_4^{3-}}{1\ \cancel{mol}\ \cancel{K_3PO_{4\,(s)}}} \qquad \boxed{[PO_4^{3-}] = 0.0325\ mol/L}$$

Dilution: - the process where additional solvent is added to lower the concentration of an original solution.

Dilution

$$C_1V_1 = C_2V_2$$

C_1 = Concentration of Original Solution **V_1 = Volume of Original Solution**

C_2 = Concentration of Diluted Solution **V_2 = Total Volume of Diluted Solution**

Example 6: Concentrated hydrochloric acid comes in 17.4 M. What is the volume of concentrated $HCl_{(aq)}$ needed to obtain 250 mL of 1.50 M of $HCl_{(aq)}$?

$C_1 = 17.4\ mol/L$

$V_1 = ?$

$C_2 = 1.50\ mol/L$

$V_2 = 250\ mL$

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1} = \frac{(1.50\ \cancel{mol/L})(250\ mL)}{(17.4\ \cancel{mol/L})}$$

$$\boxed{V_1 = 21.6\ mL}$$

Example 7: Determine the volume of water needed to dilute 30.0 mL of 0.500 M $CuSO_{4\,(aq)}$ to 0.0750 M.

C_1 = 0.500 mol/L
V_1 = 30.0 mL
C_2 = 0.0750 mol/L
V_2 = ?

$$C_1V_1 = C_2V_2$$

$$V_2 = \frac{C_1V_1}{C_2} = \frac{(0.500\ \cancel{\text{mol/L}})(30.0\ \text{mL})}{(0.0750\ \cancel{\text{mol/L}})}$$

V_2 = 200 mL

V_{water} = V_2 – 30.0 mL = ?

V_{water} = V_2 – 30.0 mL
V_{water} = 200 mL – 30.0 mL

V_{water} = 170 mL

Percent Solutions: - concentration of solution expressed in **% volume/volume** or **% mass/volume**.

Percent Solutions

$$\textbf{\% by Volume (\% (v/v))} = \frac{\textbf{Volume of Solute}}{\textbf{Total Volume of Solution}} \times \textbf{100\%}$$

(Same Unit for Both Volumes)

$$\textbf{\% by Mass (\% (m/v))} = \frac{\textbf{Mass of Solute (g)}}{\textbf{Total Volume of Solution (mL)}} \times \textbf{100\%}$$

Example 8: Calculate the percent solutions of the following.

a. 1.32 g of salt in 325 mL of water

$$\%\ (\text{m/v}) = \frac{\text{Mass of Solute (g)}}{\text{Total Volume of Solution (mL)}} \times 100\%$$

$$\%(\text{m/v}) = \frac{1.32\ \text{g salt}}{325\ \text{mL water}} \times 100\%$$

0.406% salt (m/v)

b. 45.0 mL of vinegar with 250 mL of water

$$\%\ (\text{v/v}) = \frac{\text{Volume of Solute}}{\text{Total Volume of Solution}} \times 100\%$$

$$\%\ (\text{v/v}) = \frac{45.0\ \text{mL vinegar}}{(250\ \text{mL H}_2\text{O} + 45.0\ \text{mL solute})} \times 100\%$$

15.3% vinegar (v/v)

Example 9: What is the mass of solute required to make a 3.75 L of a 22.5% sugar (m/v) solution?

$$\%\ (\text{m/v}) = \frac{\text{Mass of Solute (g)}}{\text{Total Volume of Solution (mL)}} \times 100\%$$

$$22.5\%\ (\text{m/v}) = \frac{m_{solute}}{3750\ \text{mL water}} \times 100\%$$

3.75 L = 3750 mL

$$0.225 = \frac{m_{solute}}{3750\ \text{mL water}}$$

m_{solute} = (0.225)(3750) = 843.75 g

m_{solute} = 844 g of sugar

Assignment

18.1 pg. 507 #1 and 2; pg. 508 #3 to 7
18.2 pg. 511 #8 to 11; pg. 513 #12, 13; pg. 514 #14, 15; pg. 515 #16 to 23

18.3 and 18.4: Colligative Properties of Solutions and their Calculations

Molality: - moles of solute per kilogram of solvent (m = mol/kg).

$$\textbf{Molality} = \frac{n_{solute}}{m_{solvent}}$$

Molality = *m* (mol/kg) **n = moles of Solute** **$m_{solvent}$ = Mass of Solvent in kg**

Mole Fraction (χ): - the ratio of the moles of a solute and the moles of the entire solution (solute and solvent).

Mole Fraction

$$\chi_A = \frac{n_A}{n_{Total}}$$

n_A = moles of component A **n_{Total} = Total Moles of Solution = $n_A + n_B + n_C + \ldots$**

Example 1: 40.0 g of 1-propanol with a density of 0.803 g/mL is dissolved in 150 mL of water. Describe the composition of the solution by

a. molarity b. molality c. mass percent d. mole fraction

a. Molarity

$$n = \frac{40.0\ \text{g}}{60.0962\ \text{g/mol}} = 0.6655994888\ \text{mol}\ C_3H_7OH$$

$$V_{\text{propanol}} = \frac{40.0\ \text{g}}{0.803\ \text{g/mL}} = 49.8132005\ \text{mL}$$

$$V_{\text{Total}} = 150\ \text{mL} + 49.8132005\ \text{mL}$$

$$V_{\text{Total}} = 199.8132005\ \text{mL} = 0.1998132005\ \text{L}$$

$$C_{\text{propanol}} = [C_3H_7OH] = \frac{0.6655994888\ \text{mol}}{0.1998132005\ \text{L}}$$

$$\boxed{[C_3H_7OH] = 3.33\ \text{mol/L}}$$

b. Molality

$$n = \frac{40.0\ \text{g}}{60.0962\ \text{g/mol}} = 0.6655994888\ \text{mol}\ C_3H_7OH$$

$$m_{\text{solvent}} = 150\ \text{g} = 0.150\ \text{kg}$$

(Assume 1 g ≈ 1 mL H_2O)

$$\textbf{Molality} = \frac{n_{solute}}{m_{solvent}}$$

$$\text{Molality} = \frac{0.6655994888\ \text{mol}}{0.150\ \text{kg}}$$

$$\boxed{\textbf{Molality} = 4.44\ m}$$

c. Mass Percent

$$m_{\text{propanol}} = 40.0\ \text{g}$$

$$m_{\text{solution}} = m_{\text{propanol}} + m_{\text{water}} = 40.0\ \text{g} + 150\ \text{g}$$

$$m_{\text{solution}} = 190\ \text{g}$$

$$\textbf{Mass Percent} = \frac{m_{solute}}{m_{solution}} \times 100\%$$

$$\text{Mass Percent} = \frac{40.0\ \text{g}}{190\ \text{g}} \times 100\%$$

$$\boxed{\textbf{Mass Percent} = 21.1\%}$$

d. Mole Fraction

$$n = \frac{40.0\ \text{g}}{60.0962\ \text{g/mol}} = 0.6655994888\ \text{mol}\ C_3H_7OH$$

$$n_{\text{water}} = \frac{150\ \text{g}}{18.0158\ \text{g/mol}} = 8.326024934\ \text{mol}\ H_2O$$

$$n_{\text{Total}} = n_{\text{propanol}} + n_{\text{water}}$$

$$n_{\text{Total}} = 0.6655994888\ \text{mol} + 8.326024934\ \text{mol}$$

$$n_{\text{Total}} = 8.991624422\ \text{mol}$$

$$\chi_{\text{propanol}} = \frac{n_{\text{propanol}}}{n_{Total}} = \frac{0.6655994888\ \text{mol}}{8.991624422\ \text{mol}}$$

$$\boxed{\chi_{\text{propanol}} = 0.0740}$$

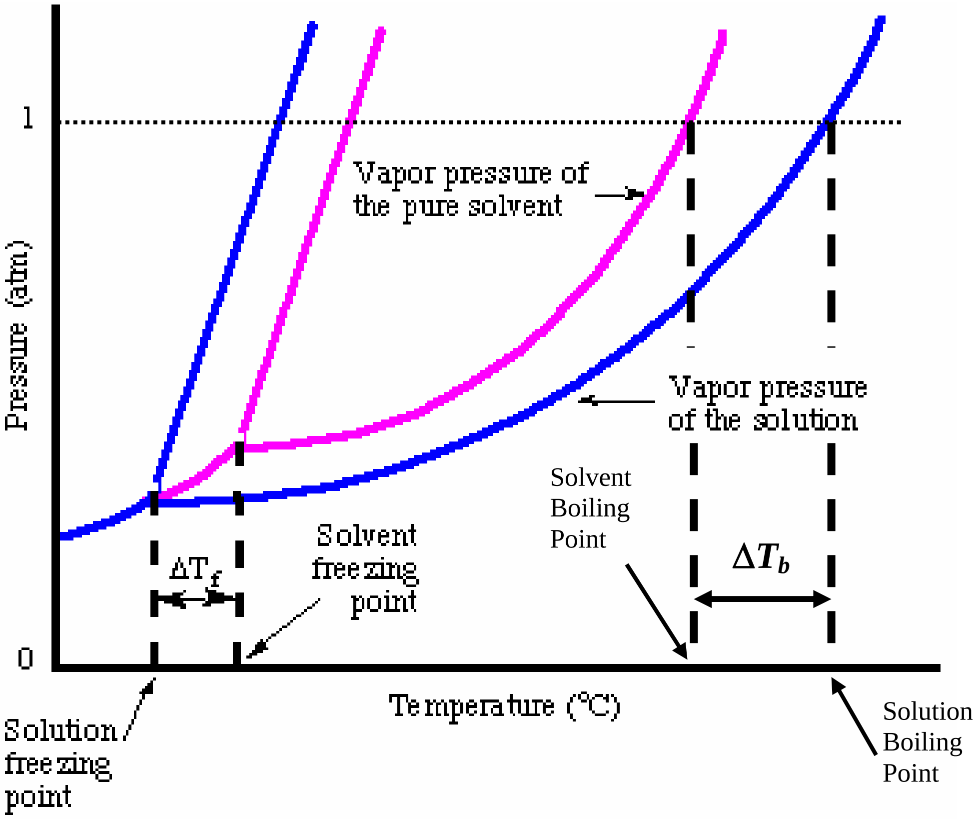
Colligative Properties: - changes in physically properties of a pure substance as it is mixed with a solute.
- these include freezing-point depression, boiling-point elevation and osmotic pressure.

Boiling-Point Elevation: - the presence of solute in a solution raises the boiling point of a pure solvent due to the extra intermolecular forces between them (more energy is needed to boil the solution as vapour pressure is lowered).
- the amount of temperature elevation to boil is directly proportional to the molality of the solute in the solution.
- can be used to **determine molar mass of a solute**.

Molal Boiling-Point Elevation Constant (K_b): - a constant relating the change in boiling point temperature and the molality of the solute in the solution.

Freezing-Point Depression: - the presence of solute decreases the vapour pressure of the solvent. Since ice has a higher vapour pressure than that of water in the solution, the freezing point has effectively been lowered. (In order for ice to form, vapour pressure of the ice has to be lowered than that of water).
- the amount of temperature depression to freeze is directly proportional to the molality of the solute in the solution.
- used most commonly to **determine molar mass of a solute**.

Molal Freezing-Point Elevation Constant (K_f): - a constant relating the change in freezing point temperature and the molality of the solute in the solution.



Boiling Point Elevation and Freezing Point Depression of Non-Electrolytic Solutions

$$\Delta T_b = K_b \times \text{Molality}_{\text{solute}} \qquad \Delta T_f = K_f \times \text{Molality}_{\text{solute}}$$

ΔT_b = Change in Boiling Point Elevation (°C)
ΔT_f = Change in Freezing Point Depression (°C)
K_b = Molal Boiling-Point Constant (°C • kg/mol)
K_f = Molal Freezing-Point Constant (°C • kg/mol)
$\text{Molality}_{\text{solute}}$ = Molality of Solute (mol/kg of solvent)

Example 2: Antifreeze, ethylene glycol (CH_2OHCH_2OH), is commonly used to prevent water from freezing in the engine in cold temperature as well as overheating. The maximum temperature a radiater can reach is 120°C. Given the densities for ethylene glycol and water are 1.11 g/mL and 1.00 g/mL respectively, and K_b and K_f for water are 0.51°C • kg/mol and 1.86 °C • kg/mol,

a. determine the volume of antifreeze needed to add to 10.0 L of water to sustain the maximum radiater temperature.

b. what is the freezing point of the solution?

a.

ΔT_b = 120°C – 100°C = 20°C
K_b = 0.51°C • kg/mol
$m_{\text{water}} = 10.0 \times 10^3$ mL × 1.00 g/mL
$m_{\text{water}} = 10.0 \times 10^3$ g = 10.0 kg
$D_{\text{antifreeze}}$ = 1.11 g/mL
$M_{\text{antifreeze}}$ = 62.0694 g/mol

$n_{\text{antifreeze}}$ = ? $m_{\text{antifreeze}}$ = ?
$V_{\text{antifreeze}}$ = ?

$$\Delta T_b = K_b \times \text{Molality}_{\text{antifreeze}} = K_b \times \frac{n_{antifreeze}}{\text{kg of water}}$$

$$n_{\text{antifreeze}} = \frac{\Delta T_b(\text{kg of water})}{K_b} = \frac{(20^\circ\,\cancel{\text{C}})(10.0\,\cancel{\text{kg}})}{(0.51^\circ\,\cancel{\text{C}} \bullet \cancel{\text{kg}}/\text{mol})} = 392.1568627 \text{ mol}$$

$$m_{\text{antifreeze}} = (392.1568627\,\cancel{\text{mol}})(62.0694\text{ g}/\cancel{\text{mol}}) = 24340.94117 \text{ g}$$

$$V_{\text{antifreeze}} = \frac{m_{antifreeze}}{D_{antifreeze}} = \frac{24340.94117\,\cancel{\text{g}}}{1.11\,\cancel{\text{g}}/\text{mL}} = 21928.77583 \text{ mL}$$

$$\boxed{V_{\text{antifreeze}} = 21.9 \text{ L}}$$

b.

K_f = 1.86°C • kg/mol
m_{water} = 10.0 kg
$n_{\text{antifreeze}}$ = 392.1568627 mol
ΔT_f = ?

$$\Delta T_f = K_f \times \text{Molality}_{\text{antifreeze}} = K_f \times \frac{n_{antifreeze}}{\text{kg of water}}$$

$$\Delta T_f = (1.86^\circ\text{C} \bullet \cancel{\text{kg}/\text{mol}})\left(\frac{392.1568627\,\cancel{\text{mol}}}{10.0\,\cancel{\text{kg}}}\right) = 72.94117646^\circ\text{C}$$

New Freezing Point = 0°C – 72.94117646°C

$$\boxed{T_{f\,(\text{soln})} = -72.9^\circ\text{C}}$$

Molar Mass Determination of a Solute using Freezing Point Depression

- one common way to determine the molar mass of a solute is dissolving it in water (if the solute is polar). After manipulating the freezing point depression formula, we can calculate molar mass of this solute.

$$\Delta T_f = K_f \times \text{Molality}_{\text{solute}}$$

$$\Delta T_f = K_f \times \frac{n_{solute}}{\text{kg of solvent}}$$

$$\frac{(\Delta T_f)(\text{kg of solvent})}{K_f} = n_{\text{solute}} \longrightarrow \frac{(\Delta T_f)(\text{kg of solvent})}{K_f} = \frac{m_{solute}}{Molar\ Mass_{solute}}$$

Molar Mass Determination from Freezing Point Depression of Non-Electrolytic Solutions

$$\text{Molar Mass}_{solute} = \frac{(K_f)(m_{solute}\text{ in grams})}{(\Delta T_f)(\text{kg of solvent})}$$

Example 3: A 20.0 g of newly synethsized enzyme can lowered the freezing point of 100.0 g CCl_4 by 4.70°C. Given that K_f for CCl_4 is 30.0°C • kg/mol, calculate the molar mass of this new enzyme.

$\Delta T_f = 4.70^\circ C$

$K_f = 30.0^\circ C \bullet kg/mol$

$m_{CCl_4} = 100.0\ g = 0.1000\ kg$

$m_{enzyme} = 20.0\ g$

$n_{enzyme} = ?$

$M_{enzyme} = ?$

$$\text{Molar Mass}_{solute} = \frac{(K_f)(m_{solute}\text{ in grams})}{(\Delta T_f)(\text{kg of solvent})}$$

$$M_{enzyme} = \frac{(30.0^\circ \cancel{C} \bullet \cancel{kg}/mol)(20.0\ g)}{(4.70^\circ \cancel{C})(0.1000\ \cancel{kg})} = 1276.595745\ g/mol$$

$$\boxed{M_{enzyme} = 1.28 \times 10^3\ g/mol}$$

Colligative Properties of Electrolytic Solutions

Because ionic solutes form electrolytes when they dissociate in solutions, we must account for the total moles of solute ions form.

van't Hoff Factor (*i*): - is a factor that relates the number of moles of all ions in a solution per mole of solutes dissolved.
- because non-electrolytic solutions do not form ions, ***i* = 1 for all non-ionic solutions**.

Colligative Properties of Electrolytic Solutions

$$\text{van't Hoff Factor } (i) = \frac{n_{ions}}{n_{solute}}$$

Boiling Point Elevation: $\Delta T_b = iK_b \times \text{Molality}_{solute}$

Freezing Point Depression: $\Delta T_f = iK_f \times \text{Molality}_{solute}$

Example 4: Determine the freezing point of a saturated salt solution (35.7 g / 100 g H_2O) if the K_f for pure water is 1.86 °C • kg/mol.

$K_f = 1.86^\circ C \bullet kg/mol$

$m_{water} = 100\ g = 0.100\ kg$

$n_{NaCl} = \dfrac{35.7\ \cancel{g}}{58.443\ \cancel{g}/mol}$

$n_{NaCl} = 0.610851599\ mol$

$i = 2$

(1 mol Na^+ and 1 mol Cl^-)

$\Delta T_f = ?$

$$\Delta T_f = iK_f \times \text{Molality}_{NaCl} = iK_f \times \frac{n_{NaCl}}{\text{kg of water}}$$

$$\Delta T_f = (2)(1.86^\circ C \bullet \cancel{kg/mol})\left(\frac{0.610851599\ \cancel{mol}}{0.100\ \cancel{kg}}\right) = 22.72367948^\circ C$$

New Freezing Point = 0°C − 22.72367948°C

$$\boxed{T_{f\ (soln)} = -22.7^\circ C}$$

Assignment

18.3 pg. 519 #24 to 27

18.4 pg. 521 #28, 29; pg. 522 #30, 31; pg. 524 #32, 33; pg. 525 #34 to 39

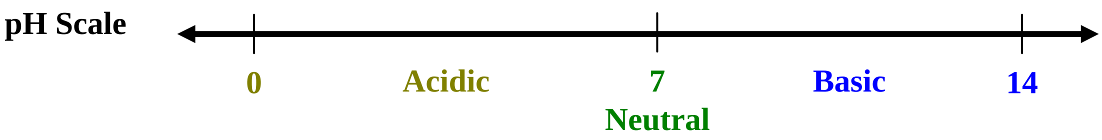
Ch 18 Review: pg. 528 #40 to 59

Chapter 20: Acids and Bases

20.1: Describing Acids and Bases

Physical and Chemical Properties of Acid and Base

Acids	**Bases**
Taste Sour (Citric Acids).	Taste Bitter.
Burning Sensation (Stomach Acid).	Feels Slippery (Detergent, Degreaser).
Corrosive with Metals (reacts to give off $H_{2\,(g)}$).	Alkaline in Nature (NaOH, Baking Soda).
Electrolytic (dissociate into ions in water)	Electrolytic (dissociate into ions in water)
Red litmus remains Red; Blue litmus turns Red.	Red litmus turns Blue; Blue litmus remains Blue.
Bromothymol Blue turns Yellow	Bromothymol Blue turns Blue.
Phenolphthalein turns Colourless.	Phenolphthalein turns Pink.
$pH < 7$	$pH > 7$



Acids: - pure substance when dissolved in aqueous solution can dissociate a H^+ ion (proton).

Bases: - pure substance when dissolved in aqueous solution can dissociate a OH^- ion (hydroxide ion).

Examples:
$HCl_{(aq)} \rightarrow \mathbf{H^+_{(aq)}} + Cl^-_{(aq)}$ ($HCl_{(aq)}$ is an Acid.)
$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + \mathbf{OH^-_{(aq)}}$ ($NaOH_{(aq)}$ is a Base.)

Nomenclature of Acid

	Ionic Compound Name		**Acid Name**
1.	**hydrogen ~ide**	→	**hydro~ic acid**
2.	**hydrogen ~ate**	→	**~ic acid**
3.	**hydrogen ~ite**	→	**~ous acid**

Example 1: Name the following acids.

a. $HI_{(aq)}$

hydrogen iod*ide*

hydro*iod*ic acid

b. $HNO_{3\,(aq)}$

hydrogen nitr*ate*

nitr*ic acid*

c. $HNO_{2\,(aq)}$

hydrogen nitr*ite*

nitr*ous acid*

Example 2: Provide chemical formulas for the following acids.

a. hydrocyanic acid

hydro*cyan*ic acid

hydrogen cyan*ide* ⇒ H^+ & CN^-

$HCN_{(aq)}$

b. oxalic acid

oxal*ic acid* ⇒ hydrogen oxal*ate*

H^+ & $OOCCOO^{2-}$

$H_2OOCCOO_{(aq)}$

c. sulfurous acid

sulfur*ous acid*

⇒ hydrogen sulf*ite*

H^+ & SO_3^{2-}

$H_2SO_{3\,(aq)}$

Nomenclature of Base: - most ionic bases come from metal cations (Group IA – Li, Na, K, Rb, Cs and some of Group IIA – Ca, Ba, Sr) with OH^- ion.

- the naming of these ionic bases are "metal name" hydroxide.

Example 3: Name the following bases.

a. $KOH_{(aq)}$

K^+ = potassium

OH^- = hydroxide

potassium hydroxide

b. $Ba(OH)_{2\,(aq)}$

Ba^{2+} = barium

OH^- = hydroxide

barium hydroxide

Example 4: Provide chemical formulas for the following acids.

a. strontium hydroxide

Sr^{2+} and OH^- ⇒ **$Sr(OH)_2$**

Need 1 Sr^{2+} & 2 (OH^-) to balance

b. rubidium hydroxide

Rb^+ and OH^- ⇒ **$RbOH$**

Need 1 Rb^+ & 1 OH^- to balance

20.2: Hydrogen Ions and Acidity

Conceptual Definition: - an explanation that attempts to describe why things are the way they are.

Arrhenius Concept: - **acids are H^+ (proton) producers** and **bases are OH^- producers**.

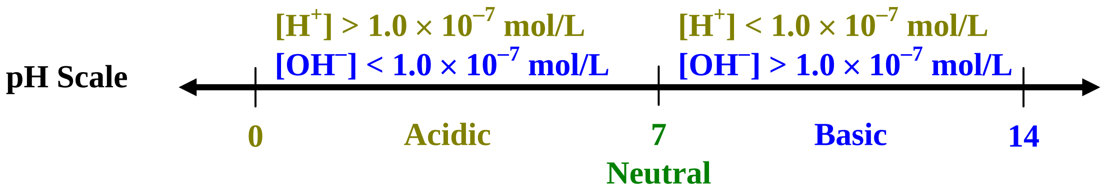
Examples:

$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ ($HCl_{(aq)}$ is an Arrhenius Acid.)

$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$ ($NaOH_{(aq)}$ is an Arrhenius Base.)

pH Scale: - a logarithmic scale to **measure the acidity (relative $[H^+]$)** of a solution.

- the **lower the pH, the more acidic** (less basic) is the solution (**more $[H^+]$ and less $[OH^-]$**).
- the **higher the pH, the more basic** (less acidic) is the solution (**less $[H^+]$ and more $[OH^-]$**).
- it is normally reported between 0 to 14 (**with 7 as neutral**), but it **can be above 14 (very basic) or below 0 (very acidic).**
- **an increase of 1 on a pH scale means a decrease of $[H^+]$ by a factor of 10; an increase of 2 on a pH scale means a decrease of $[H^+]$ by a factor of 100.**



Acidic Solution: - solutions that has a **pH < 7**, where **$[H^+] > 1.0 \times 10^{-7}$ M** and **$[OH^-] < 1.0 \times 10^{-7}$ M**.

Basic Solution: - solutions that has a **pH > 7**, where **$[H^+] < 1.0 \times 10^{-7}$ M** and **$[OH^-] > 1.0 \times 10^{-7}$ M**.
- also known as **Alkaline Solution** (basic solutions that were from alkali metal hydroxides (LiOH, NaOH, KOH, RbOH, and CsOH) or alkaline metal hydroxides ($Ca(OH)_2$, $Ba(OH)_2$, and $Sr(OH)_2$)

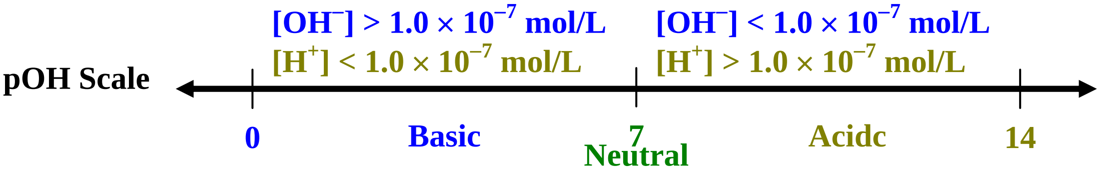
Example: pH of Some Common Substances

Substance	pH
1 M of HCl	0.00
Stomach Acid	2.00
Lemon Juice	2.50
Vinegar	3.00

Substance	pH
Milk	6.30
Rain Water	6.70
Pure Water	7.00
Blood	7.50

Substance	pH
1 M of Baking Soda ($NaHCO_3$)	9.68
Ammonia as Household Cleaner	12.00
1 M of NaOH	14.00

pOH Scale: - a logarithmic scale to **measure the basicity (relative $[OH^-]$)** of a solution.
- the **lower the pOH, the more basic** (less acidic) is the solution (**more $[OH^-]$ and less $[H^+]$**).
- the **higher the pOH, the less basic** (more acidic) is the solution (**less $[OH^-]$ and more $[H^+]$**).
- it is normally reported between 0 to 14 (**with 7 as neutral**), but it **can be above 14 (very acidic) or below 0 (very basic).**
- **an increase of 2 on a pOH scale means a decrease of $[OH^-]$ by a factor of 10; an increase of 2 on a pOH scale means a decrease of $[OH^-]$ by a factor of 100.**



pH and pOH Scales

$$pH = -\log [H^+] \qquad pOH = -\log [OH^-]$$

$$pH + pOH = 14.00$$

Example 1: Calculate the pH and the pOH for the following solutions.

a. $[H^+] = 0.0100$ mol/L

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

$pH = -\log(0.0100)$

$\boxed{pH = 2.00}$

$pOH = 14 - pH$

$pOH = 14 - 2.00$

$\boxed{pOH = 12.00}$

b. $[H^+] = 5.00 \times 10^{-6}$ M

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

$pH = -\log(5.00 \times 10^{-6})$

$\boxed{pH = 5.30}$

$pOH = 14 - pH$

$pOH = 14 - 5.30$

$\boxed{pOH = 8.70}$

c. 1.50 M of $KOH_{(aq)}$

$$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$$

(Ionic Strong Base: $[KOH]_0 = [OH^-] = 1.50$ M)

$pOH = -\log [OH^-]$

$pOH = -\log(1.50)$

$\boxed{pOH = -0.176}$

$pH = 14 - pOH$

$pH = 14 - (-0.176)$

$\boxed{pH = 14.176}$

d. 1.00×10^{-3} mol/L of $Ba(OH)_{2\,(aq)}$_

$$Ba(OH)_{2\,(aq)} \rightarrow Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

(Ionic Strong Base:

$[OH^-] = 2[Ba(OH)_2] = 2.00 \times 10^{-3}$ M)

$pOH = -\log [OH^-]$

$pOH = -\log(2.00 \times 10^{-3})$

$\boxed{pOH = 2.70}$

$pH = 14 - pOH$

$pH = 14 - (2.70)$

$\boxed{pH = 11.30}$

Example 2: Calculate the $[H^+]$ and the $[OH^-]$ for the following solutions.

a. pH = 4.00

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

$[H^+] = 10^{-pH} = 10^{-4.00}$

$\mathbf{[H^+] = 1.00 \times 10^{-4}\ M}$

$pOH = 14 - pH$

$pOH = 14 - 4.00 = 10.00$

$pOH = -\log [OH^-]$

$[OH^-] = 10^{-pOH} = 10^{-10.00}$

$\mathbf{[OH^-] = 1.00 \times 10^{-10}\ M}$

b. pOH = 3.00

$pOH = -\log [OH^-]$

$[OH^-] = 10^{-pOH} = 10^{-3.00}$

$\mathbf{[OH^-] = 1.00 \times 10^{-3}\ M}$

$pH = 14 - pOH$

$pH = 14 - 3 = 11.00$

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

$[H^+] = 10^{-pH} = 10^{-11.00}$

$\mathbf{[H^+] = 1.00 \times 10^{-11}\ M}$

Acid-Base Indicators:
- chemicals that change colours at a specific pH range.
- they are themselves organic acids. Since they are usually very big structurally, we use abbreviations to describe them in chemical equations.
- the acidic form of the indictor has a different colour than its basic form.

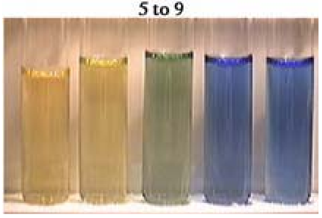
Example: Bromothymol Blue ($HBb_{(aq)}$ /$Bb^-_{(aq)}$) changes colours from yellow to blue at a pH range of 6.0 to 7.6. At its **mid-range (6.8 – average of 6.0 and 7.6)**, the colour would be **green (yellow + blue)**.

$$\underset{\text{(Yellow)}}{\mathbf{HBb_{(aq)}}} \rightleftharpoons H^+_{(aq)} + \underset{\text{(Blue)}}{\mathbf{Bb^-_{(aq)}}}$$

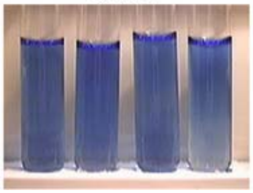
At pH ≤ 6, bromothymol blue is yellow. At pH ≥ 8, bromothymol blue is blue.

1 to 4

5 to 9



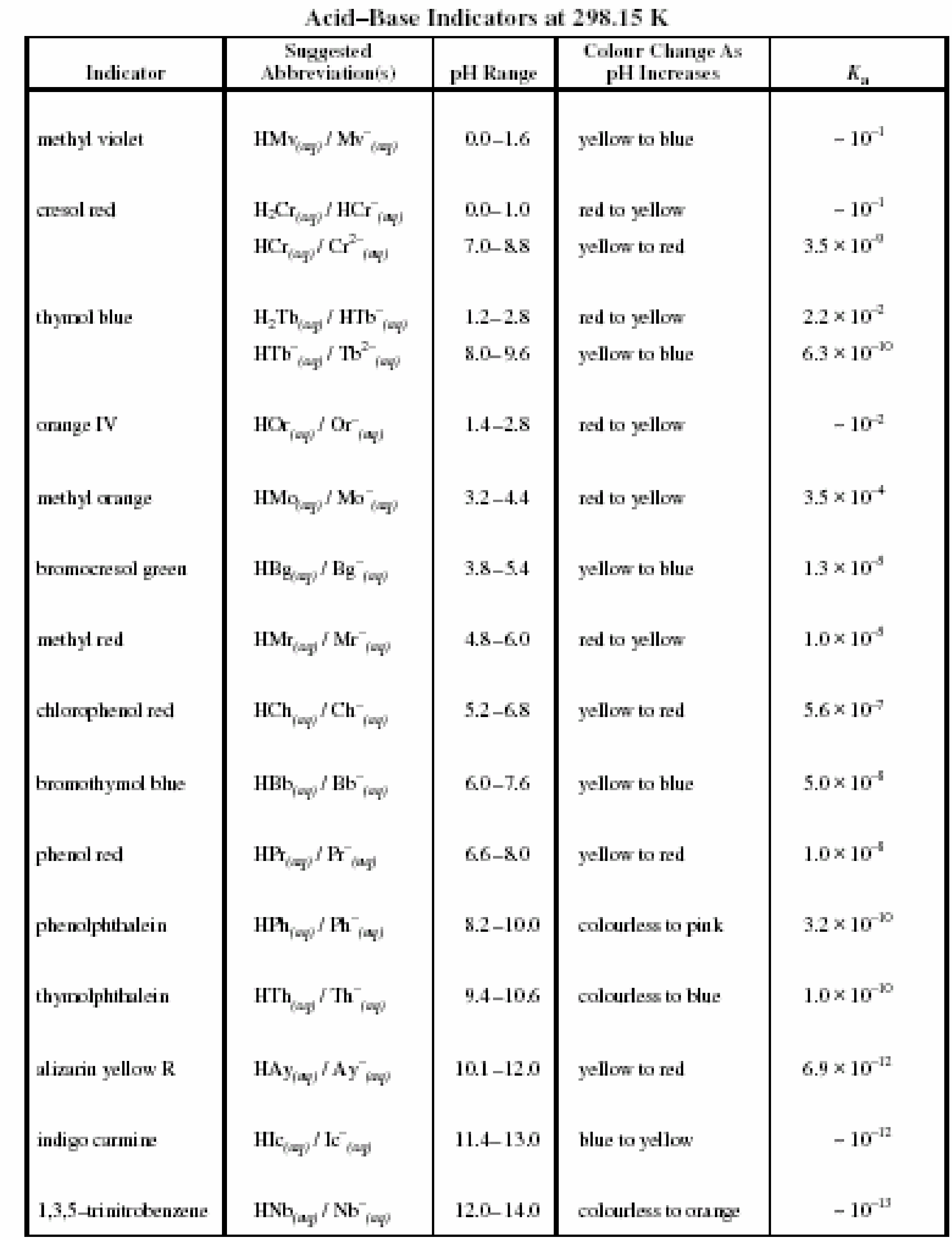
10 to 13



pH Meter: - a device consists of an electrode that accurately measures pH.

Example 3: Use the indicator table on the next page, complete the following table for a solution has a pH of 4.5.

Indicators	Observed Colours	Indicators	Observed Colours
methyl violet	**blue**	bromocresol green	**green**
methyl orange	**yellow**	methyl red	**red**
phenolphthalein	**colourless**	bromothymol blue	**yellow**



Acid–Base Indicators at 298.15 K

Indicator	Suggested Abbreviation(s)	pH Range	Colour Change As pH Increases	K_a
methyl violet	$HMv_{(aq)}$ / $Mv^-_{(aq)}$	0.0–1.6	yellow to blue	$\sim 10^{-1}$
cresol red	$H_2Cr_{(aq)}$ / $HCr^-_{(aq)}$ $HCr^-_{(aq)}$ / $Cr^{2-}_{(aq)}$	0.0–1.0 7.0–8.8	red to yellow yellow to red	$\sim 10^{-1}$ 3.5×10^{-9}
thymol blue	$H_2Tb_{(aq)}$ / $HTb^-_{(aq)}$ $HTb^-_{(aq)}$ / $Tb^{2-}_{(aq)}$	1.2–2.8 8.0–9.6	red to yellow yellow to blue	2.2×10^{-2} 6.3×10^{-10}
orange IV	$HOr_{(aq)}$ / $Or^-_{(aq)}$	1.4–2.8	red to yellow	$\sim 10^{-2}$
methyl orange	$HMo_{(aq)}$ / $Mo^-_{(aq)}$	3.2–4.4	red to yellow	3.5×10^{-4}
bromocresol green	$HBg_{(aq)}$ / $Bg^-_{(aq)}$	3.8–5.4	yellow to blue	1.3×10^{-5}
methyl red	$HMr_{(aq)}$ / $Mr^-_{(aq)}$	4.8–6.0	red to yellow	1.0×10^{-5}
chlorophenol red	$HCh_{(aq)}$ / $Ch^-_{(aq)}$	5.2–6.8	yellow to red	5.6×10^{-7}
bromothymol blue	$HBb_{(aq)}$ / $Bb^-_{(aq)}$	6.0–7.6	yellow to blue	5.0×10^{-8}
phenol red	$HPr_{(aq)}$ / $Pr^-_{(aq)}$	6.6–8.0	yellow to red	1.0×10^{-8}
phenolphthalein	$HPh_{(aq)}$ / $Ph^-_{(aq)}$	8.2–10.0	colourless to pink	3.2×10^{-10}
thymolphthalein	$HTh_{(aq)}$ / $Th^-_{(aq)}$	9.4–10.6	colourless to blue	1.0×10^{-10}
alizarin yellow R	$HAy_{(aq)}$ / $Ay^-_{(aq)}$	10.1–12.0	yellow to red	6.9×10^{-12}
indigo carmine	$HIc_{(aq)}$ / $Ic^-_{(aq)}$	11.4–13.0	blue to yellow	$\sim 10^{-12}$
1,3,5-trinitrobenzene	$HNb_{(aq)}$ / $Nb^-_{(aq)}$	12.0–14.0	colourless to orange	$\sim 10^{-13}$

Assignment

20.1: pg. 579 #1 to 5; pg. 609 #34

20.2: pg. 582 #7; pg. 586 #8 and 9; pg. 587 #10 and 11; pg. 588 #12 and 13; pg. 589 #14 and 15; pg. 593 #16 to 18; pg. 609 #36 to 45

21.1: Neutralization

Neutralization: - the reaction of an acid and a base to form water and an acid-base salt.

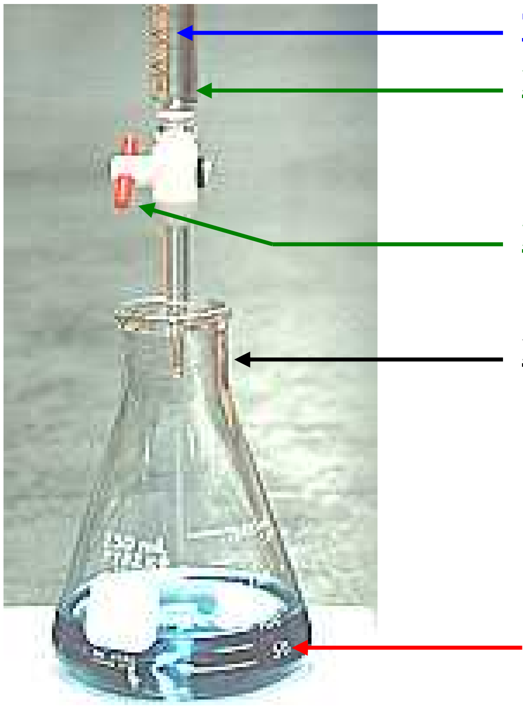
Acid-Base Salt: - simply known as **salt** – mostly soluble ionic compound left after from a neutralization reaction.
- for strong acid and strong base neutralization, it consists of the anions of a strong acid and a cations of a strong base.

Different Forms of Strong Acid and Strong Base Neutralization Chemical Equation.

1. **Molecular Equation:** $HA_{(aq)} + MOH_{(aq)} \rightarrow HOH_{(l)} + \boxed{MA_{(aq)}} \longleftarrow \text{Acid-Base Salt}$
2. **Complete Ionic Equation:** $H^+_{(aq)} + A^-_{(aq)} + M^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + M^+_{(aq)} + A^-_{(aq)}$
3. **Net-Ionic Equation:** $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Titration: - a volumetric analysis that involves measuring the volume of known concentration solution to measure a given volume of an unknown concentration solution.

Titration Set-up



Titrant: - the solution of known concentration.

Buret: - a precise apparatus to deliver the titrant.
- the volume of the titrant added is read by subtracting the final volume and the initial volume.

Buret Valve: - can be adjusted to let one drop out at a time.

Erlenmeyer Flask: - a container commonly uses to hold the analyte. (Narrow mouth prevents splash and spillage.)

Analyte: - the solution of an unknown concentration.
- the exact volume is usually delivered by a pipet.

Acid-Base Titration: - volumetric analysis that assist in determining the unknown concentration in an acid and base neutralization.

Equivalent Point (Stoichiometric Point): - a point where the number of moles of H^+ is equivalent to the number of moles of OH^-. ($n_{H^+} = n_{OH^-}$)

Endpoint: - a point where the indicator actually changes colour to indicate neutralization is completed.

Indicator: - a chemical that changes colour due to the pH of the solution (more in the next section).

Example 1: Use the following observation table to determine the concentration of nitric acid.

10.0 mL of $HNO_{3\,(aq)}$ titrated by 0.0350 mol/L of $NaOH_{(aq)}$

	Trial 1	**Trial 2**	**Trial 3**	**Trial 4**
Initial Volume	0.32 mL	24.19 mL	3.48 mL	24.97 mL
Final Volume	24.19 mL	45.71 mL	24.97 mL	46.47 mL
Volume of NaOH added	**23.87 mL**	**21.52 mL**	**21.49 mL**	**21.50 mL**
Bromothymol Blue Colour	Blue	Green	Green	Green

First, we have to complete the table by subtracting the final and the initial volumes. Since the titration is completed when the indicator turns green, we only average the result of the last 3 trials.

Average Volume of NaOH added = $\frac{21.52\text{ mL} + 21.49\text{ mL} + 21.50\text{ mL}}{3}$ **= 21.50 mL**

$NaOH_{(aq)}$	+	$HNO_{3\,(aq)}$	→	$HOH_{(l)}$	+	$NaNO_{3\,(aq)}$
21.50 mL		10.0 mL				
0.0350 mol/L		? mol/L				

① $n_{\text{NaOH}} = CV = (0.0350\text{ mol/}\cancel{\text{L}})\,(21.50\text{ m}\cancel{\text{L}}) = 0.7525\text{ mmol}$

② $n_{\text{HNO}_3} = 0.7525\text{ m}\cancel{\text{mol NaOH}} \times \frac{1\text{ mol HNO}_3}{1\ \cancel{\text{mol NaOH}}} = 0.7525\text{ mmol HNO}_3$

③ $C_{\text{HNO}_3} = \frac{n}{V} = \frac{0.7525\ \cancel{\text{m}}\text{mol}}{10.0\ \cancel{\text{m}}\text{L}} = 0.07525\text{ mol/L}$

$[HNO_3]$ = 0.0753 mol/L = 75.3 mmol/L

Example 2: Calculate the concentration of 10.00 mL of $Na_2S_{(aq)}$ when it is completely neutralized with 27.88 mL.of 0.0350 M of $HBr_{(aq)}$.

We have to write out the balance chemical equation first.

$2\ HBr_{(aq)}$	+	$Na_2S_{(aq)}$	→	$H_2S_{(aq)}$	+	$2\ NaBr_{(aq)}$
27.88 mL		10.0 mL				
0.0350 mol/L		? mol/L				

① $n_{\text{HBr}} = CV = (0.0350\text{ mol/}\cancel{\text{L}})\,(27.88\text{ m}\cancel{\text{L}}) = 0.9758\text{ mmol}$

② $n_{\text{Na}_2\text{S}} = 0.9758\text{ m}\cancel{\text{mol HBr}} \times \frac{1\text{ mol Na}_2\text{S}}{2\ \cancel{\text{mol HBr}}} = 0.4879\text{ mmol Na}_2\text{S}$

③ $[Na_2S] = \frac{n}{V} = \frac{0.4879\ \cancel{\text{m}}\text{mol}}{10.0\ \cancel{\text{m}}\text{L}}$

$[Na_2S]$ = 0.0488 mol/L

Assignment

21.1 pg. 616 #1 and 2; pg. 618 #3 and 4; pg. 624 #17 to 19;
pg. 640 #36 to 39; pg. 641 # 62