Chem 20 Final Review

CHEMICAL BONDING

Bonding Theory: Valence Electrons & Orbitals

- According to bonding theory, valence electrons are classified in terms of *orbital occupancy*.

  (0 = empty, 1 = half filled , 2 = full)

  - An atom with a valence orbital that has a single electron can theoretically share that electron with another atom
    - Such an electron is called a **BONDING ELECTRON**

  - An atom with a full valence orbital (2 e^-’s), repels nearby orbitals and wants to be alone
    - Such a pairing is called a **LONE PAIR**

Covalent Bonding

- Both atoms have a high EN so neither atom "wins"
  - The simultaneous attraction of two nuclei for a shared pair of bonding electrons = covalent bond
  - EN difference can be zero = Cl – Cl

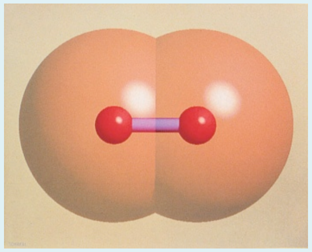
    EN = 3.2 EN = 3.2

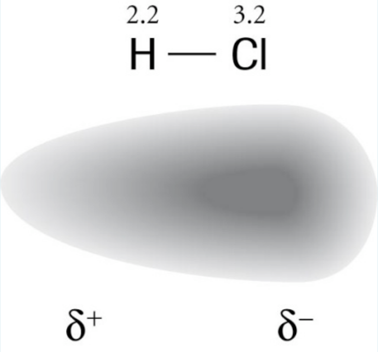
  - EN difference can be small = H - Cl

    EN = 2.2 EN = 3.2

    - *This is called a polar covalent bond – because one side pulls on the electrons more but we will learn more about this in Section 3.3*

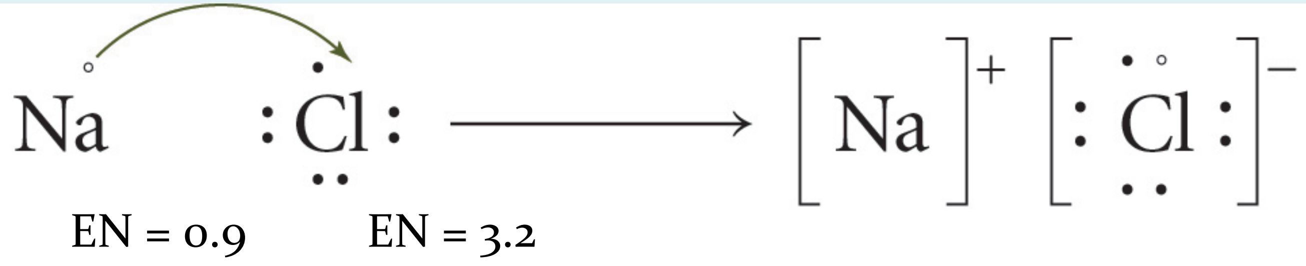
Cl_2 = diatomic

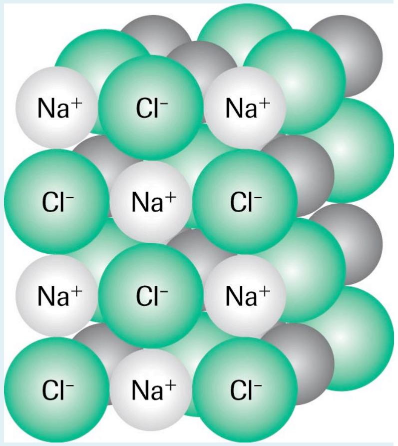




Ionic Bonding

- The EN of the two atoms are quite different
- The atom with the higher EN will remove the bonding e^- from the other atom
- Electron transfer occurs
  - Positive and negative ions are formed which electrically attract each other





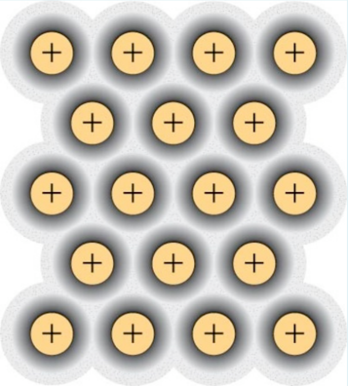
Metallic Bonding

- Both atoms have a relatively low EN so atoms share valence electrons, but no actual chemical reaction takes place

- In metallic bonding:
  a) e^-'s are not held very strongly by their atoms
  b) the atoms have vacant valence orbitals

  \- *This means the electrons are free to move around between the atoms and the (+) nuclei on either side will attract them*

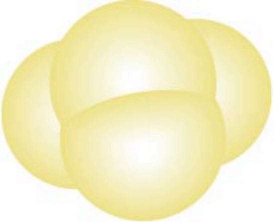
Analogy: *The positive nuclei are held together by a glue of negative e-'s*

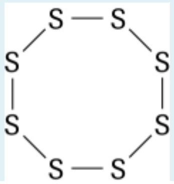


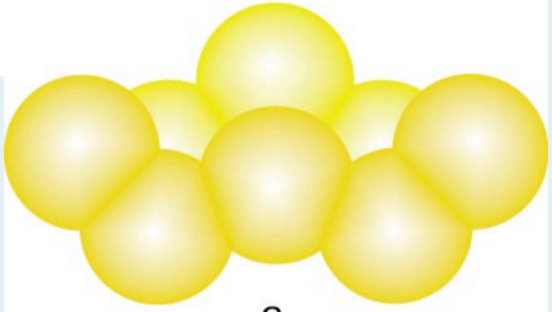
Molecular Elements

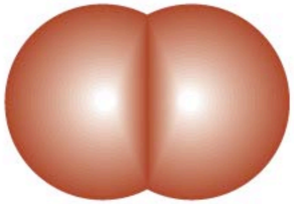
- Many molecular elements are *diatomic* and some are *polyatomic*
  - You will need to memorize the formulas of the 9 molecular elements as they will not be given to you:


H_2


P_4



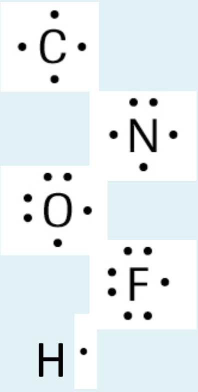

S_8


Br_2

Name	Symbol
hydrogen	$H_2(g)$
nitrogen	$N_2(g)$
oxygen	$O_2(g)$
fluorine	$F_2(g)$
chlorine	$Cl_2(g)$
iodine	$I_2(g)$
bromine	$Br_2(g)$
phosphorous	$P_4(g)$
sulfur	$S_8(g)$

Determining Lewis Formulas

- So why do we care about **bonding capacity?**
  - If we know how many ***bonding e^-'s*** an atom has, we can predict what structure a molecular compound



Atom	Number of valence electrons	Number of bonding electrons	Bonding capacity
carbon	4	4	4
nitrogen	5	3	3
oxygen	6	2	2
halogens	7	1	1
hydrogen	1	1	1

I.e. Carbon can form 4 single bonds, 2 double bonds, 1 triple and 1 single, or 1 double and 2 singles

Lewis Formulas – Guided Ex. #2

Determine the Lewis formula & structural formula for the nitrate ion, NO_3^-

1. Count the valence electrons (*look for a net charge if an ion).

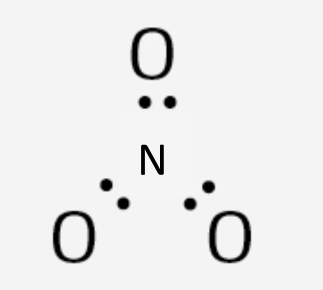
nitrogen = 1 x 5 valence e-'s = 5

oxygen = 3 x 6 valence e-'s = 18

23 + 1 (b/c net charge is -1) = 24

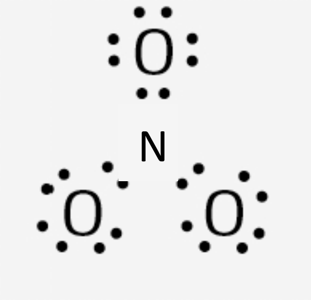
2. Which is the central atom? Nitrogen (in lesser quantity)

3. Arrange peripheral atoms around central atom and place 1 pair of valence e-'s between them



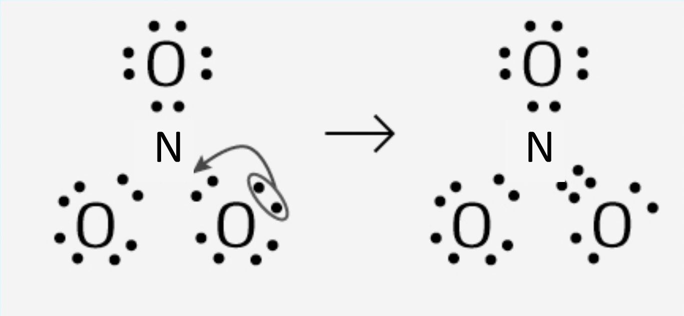
Lewis Formulas – Guided Ex. #2

4. Place lone pairs on all peripheral atoms to complete their octet

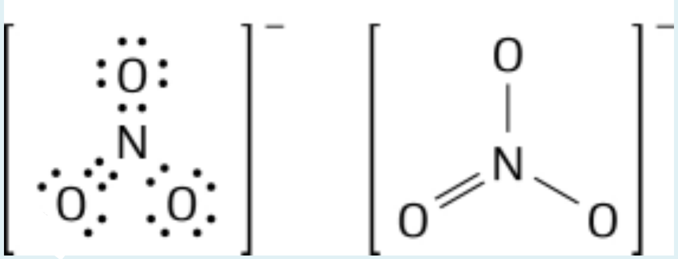


5. Place any remaining valence e-'s on the central atom as lone pairs.

6. If the central atom's octet is not complete, move a lone pair from a peripheral atom to a new position between the peripheral and central atom.



7. If the entity is a **polyatomic ion**, place square brackets around the entire Lewis formula and then write the **net charge** outside the bracket on the upper right.



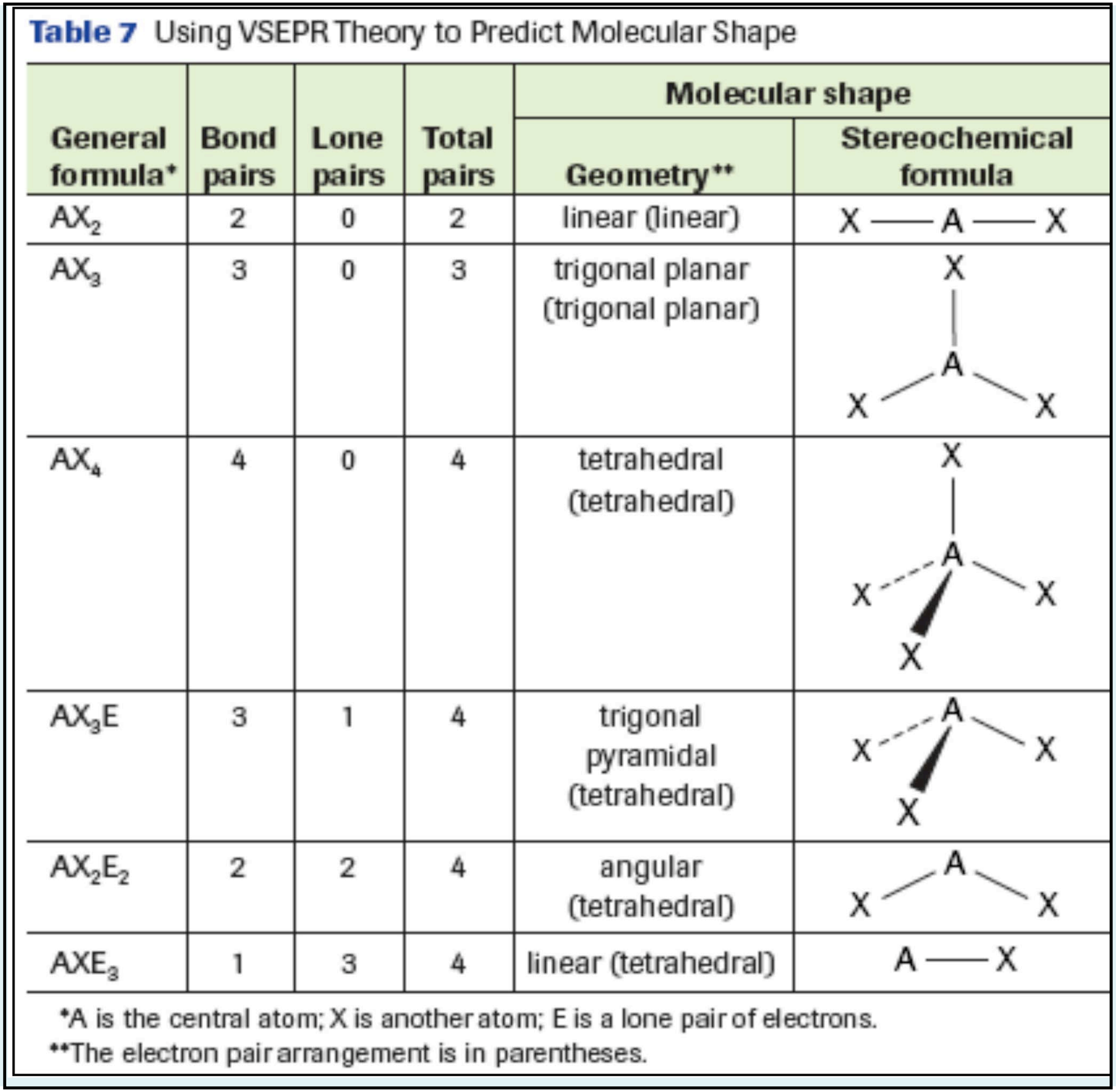


Table 7 Using VSEPR Theory to Predict Molecular Shape

General formula*	Bond pairs	Lone pairs	Total pairs	Molecular shape	
				Geometry**	Stereochemical formula
AX_2	2	0	2	linear (linear)	X — A — X
AX_3	3	0	3	trigonal planar (trigonal planar)	
AX_4	4	0	4	tetrahedral (tetrahedral)	
AX_3E	3	1	4	trigonal pyramidal (tetrahedral)	
AX_2E_2	2	2	4	angular (tetrahedral)	
AXE_3	1	3	4	linear (tetrahedral)	A — X

*A is the central atom; X is another atom; E is a lone pair of electrons.

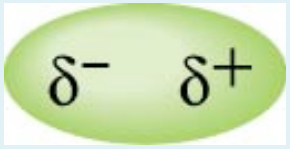
**The electron pair arrangement is in parentheses.

Pg. 95

Polarity

- Chemists believe that molecules are made up of charged particles (electrons and nuclei).

- A **polar** molecule is one in which the negative (electron) charge is not distributed symmetrically among the atoms making up the molecule.



  - Thus, it will have partial positive and negative charges on opposite sides of the molecule.

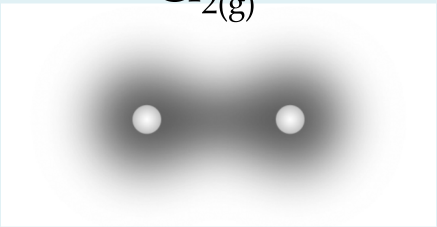
- A molecule with symmetrical electron distribution is a **nonpolar** molecule.

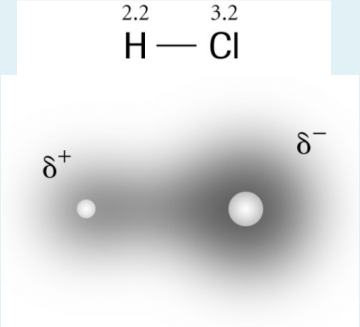
- The existence of polar molecules can be demonstrated by running a stream of water past a charged object.
  - *Demo: See Figure 9*

PREDICTING AND EXPLAINING POLARITY

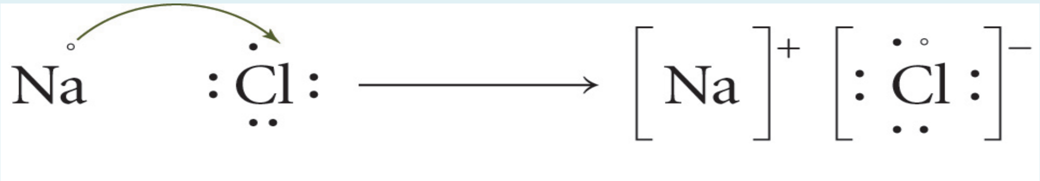
- Pauling explained the polarity of a **covalent bond** as the difference in electronegativity of the bonded atoms.
  - If the bonded atoms have the same electronegativity, they will attract any shared electrons equally and form a **nonpolar covalent bond**.
  - If the atoms have different electronegativities, they will form a **polar covalent bond**.
  - The greater the electronegativity difference, the more polar the bond will be.
- For a very large electronegativity difference, the difference in attraction may transfer one or more electrons resulting in **ionic bonding**.

$Cl_{2(g)}$





We use the Greek symbol delta to show partial charges

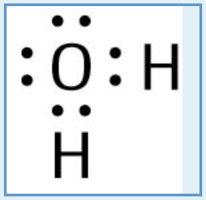


Guided Practice #1

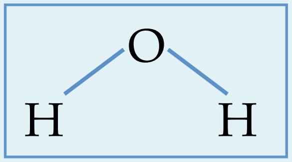
Go to Learning Tip pg. 102

- Predict the polarity of the water molecule.

1) Draw the Lewis formula



2) VSEPR: Draw the stereochemical formula

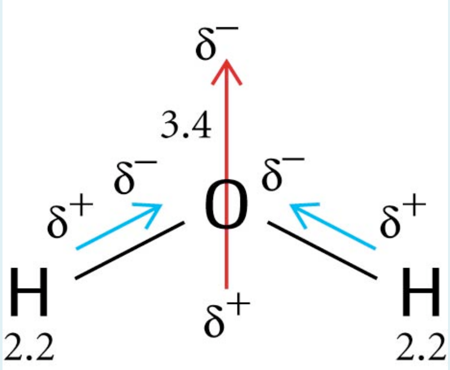


Angular (bent)

3) Assign the EN of the atoms, assign $\delta -$ and $\delta +$ to the bonds

4)

- The bond dipoles (vectors) do not balance.
- Instead, they add together to produce a nonzero molecular dipole (shown in red).
- This results in a polar molecule (explains bending water)



BACKGROUND

- There are three types of forces in matter:

1) Intranuclear force (bond) – bonds within the nucleus between protons and neutrons (very strong)

2) Intramolecular force (bond) – bonds between atoms **within** the molecule or between ions **within** the crystal lattice (quite strong)

3) Intermolecular force (bond) – bonds **between** molecules (quite weak); are electrostatic (involve positive and negative charges)

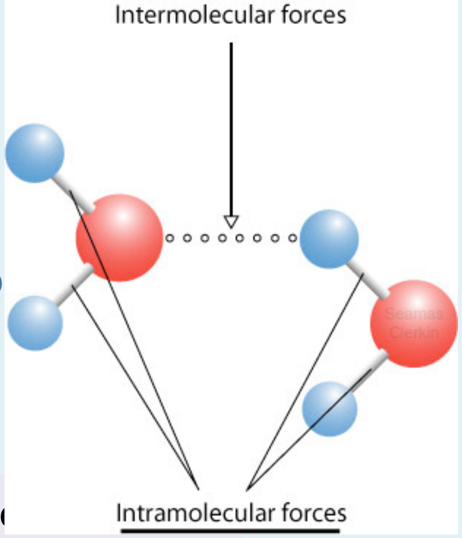
There are 3 types of intermolecular bonds:

Weakest a) Dipole-Dipole Forces (a.k.a. Polar Forces)

Medium b) London Force (a.k.a. London Dispersion Force, Dispersio

Strongest c) Hydrogen Bonding

Note: "Van der Walls force" – includes London and dipole-dipole forc



1) Dipole-Dipole Force

- *The simultaneous attraction between oppositely charged ends of polar molecules.*
  - Simply put, the attraction between diploes

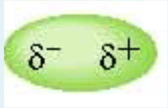
$\overset{\delta+}{H}-\overset{\delta-}{Cl}\cdots\overset{\delta+}{H}-\overset{\delta-}{Cl}$

Dipole: a partial separation of positive and negative charges within a molecule, due to electronegativity differences

- Dipole-dipole forces are among the **weakest** intermolecular forces, but still control important properties (i.e. Solubility because water is polar))

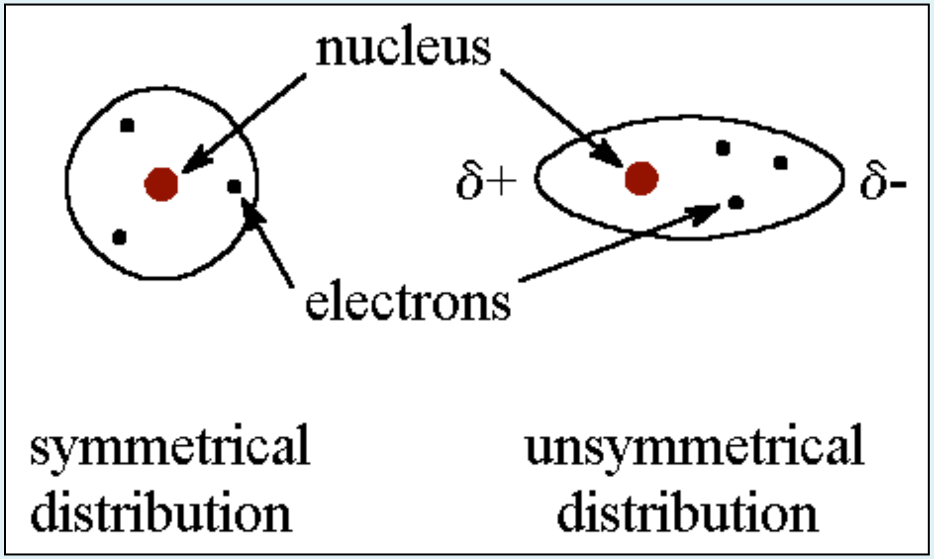
2) London Force

- *Simultaneous attraction between a **momentary dipole** in a molecule and the momentary dipoles in surrounding molecules*



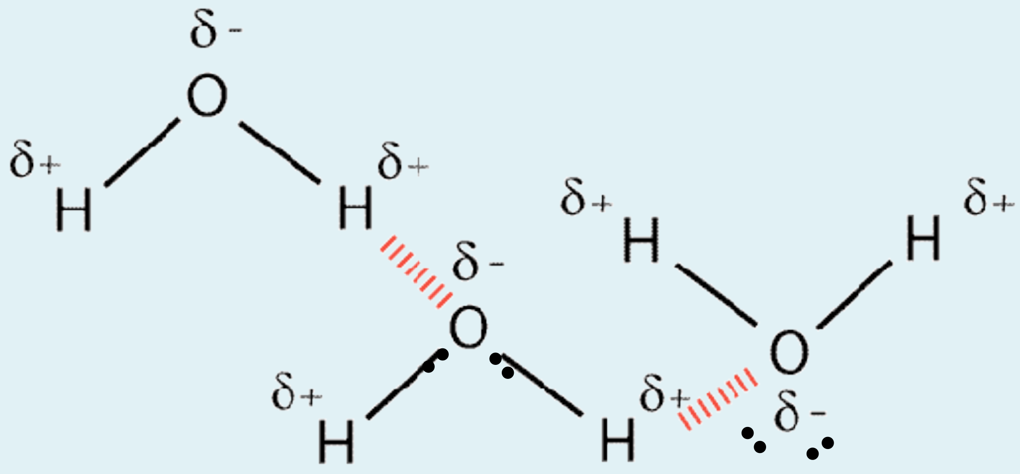
momentary dipole: an uneven distribution of electrons around a molecule, resulting in a temporary charge difference between its ends

They last for just the instant that the electrons are not distributed perfectly even.



3) **Hydrogen Bonding**

- Occurs when a hydrogen atom bonded to a strongly electronegative atom, (N, O and F) is attracted to a lone pair of electrons in an adjacent molecule.
  - Hydrogen nucleus (proton) is simultaneously attracted to two pairs of electrons; one closer (in the same molecule) and one further away (on the next molecule)



Why do you need a strongly electronegative atom?

It pulls the hydrogen's electron away making it "unshielded", so the lone pair on the other side can come much closer

GASES

Comparing Kelvin and Celsius Scales

- To convert degrees Celsius to Kelvin , you add 273.

  K = °C + 273

- To convert Kelvin to degrees Celsius, you subtract 273.

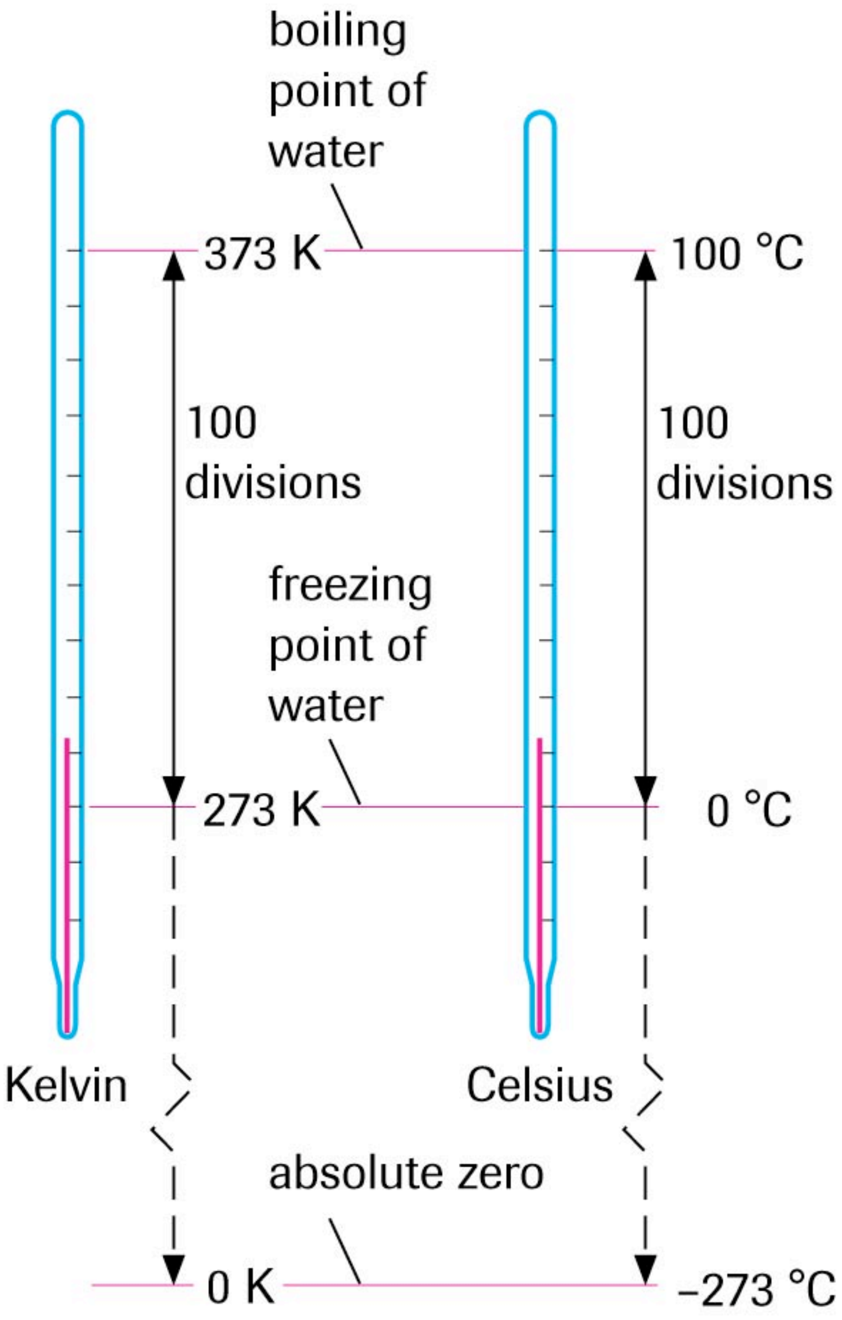
  °C = K - 273

- Examples:
  - What is 254 K in °C ?

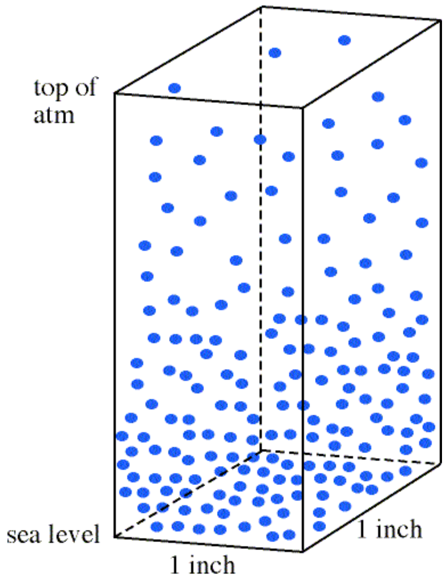
    -19°C

  - What is -34°C in K ?

    239K



Atmospheric Pressure



- *Pressure exerted by air on all objects*

Standard Temperature and Pressure (STP)

= 101.325 kPa and 0°C

- But laboratory temperatures are not at 0°C
- So scientists agreed on another set of conditions...

Standard Ambient Temperature and Pressure (SATP)

= 100 kPa and 25°C

- Much closer to lab conditions – so scientists don't **freeze**

Molar mass

- Now you try it…

- What is the molar mass of methanol CH_3OH?

C = 12.01 g/mol x 1 = 12.01 g/mol

H = 1.01g/mol x 4 = 4.04 g/mol

O = 16.00g/mol x 1 = $\underline{\text{16.00 g/mol}}$

Molar mass of methanol = 32.05 g/mol

Gas Laws

- **They are based on the temperature, pressure and volume relationships that all gases have in common**

1. **Boyle's Law** $P_1V_1 = P_2V_2$
2. **Charles' Law** $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
3. **Combined Gas Law** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Example: Law of Combining Volumes

Use the law of combining volumes to predict the volume of oxygen required for the complete combustion of 120 mL of butane gas from a lighter.

1) The first step is to write the balanced chemical equation, including what you are given and what you need to find:

$$\underset{120\ \text{mL}}{\mathbf{2}C_4H_{10}(g)} + \underset{V = ?}{\mathbf{13}O_2(g)} \rightarrow \mathbf{8}CO_2(g) + \mathbf{10}H_2O(g)$$

2) From this chemical equation you can see that 13 mol of oxygen is required for every 2 mol of butane. Therefore, the volume of oxygen has to be greater than 120mL by a factor of 13/2.

$$V_{O_2}:\quad 120\ \cancel{\text{ml}\ C_4H_{10}}\ \text{x}\ \left(\frac{13\ \text{mL}\ O_2}{2\ \cancel{\text{ml}\ C_4H_{10}}}\right) = 780\ \text{mL}$$

To make sure that the ratio is used in the correct order, you could include the chemical formula with each quantity as shown above. Note the cancellation of the units and chemical formulas

Molar Volume

STP = 22.4L/mol
SATP = 24.8 L/mol

- Molar volume is the **same** for all gases at the same temperature and pressure (remember, all gases have the same physical properties)
  - At STP, molar volume = 22.4 L/mol (101.325 kPa and 0°C)
  - At SATP, molar volume = 24.8 L/mol (100 kPa and 25°C)
    - This can be used as a conversion factor just like molar mass!

At STP, one mole of gas has a volume of 22.4 L, which is approximately the volume of 11 "empty" 2 L pop bottles.

Molar Volume – Practice

4. A propane tank for a barbecue contains liquefied propane. IF the tank mass drops by 9.1 kg after a month's use, what volume of propane gas at SATP was used for cooking?

Molar mass (M): $C_3H_{8(g)}$ = 44.11 g/mol

STP = 22.4L/mol
SATP = 24.8 L/mol

$$V_{C_3H_{8(g)}}:\ 9.1\ \text{k}\cancel{\text{g}} \times \left(\frac{1\ \cancel{\text{mol}}}{44.11\cancel{\text{g}}}\right) \times \left(\frac{24.8\ \text{L}}{1\ \cancel{\text{mol}}}\right) = 5.1\ \text{kL}$$

- What if I wanted your answer in litres?

$$5.1\ \cancel{\text{kL}} \times \frac{1000\text{L}}{1\ \cancel{\text{kL}}} = 5100\ \text{L} = 5.1 \times 10^3\ \text{L}$$

IDEAL GAS LAW

- Before getting too far into this law, it is important to understand the difference between an ideal gas and a real gas....
  - IDEAL GAS – does not really exist, it is hypothetical
    - Follows all gas laws perfectly under all conditions
    - Does not condense when cooled
    - Assumes that the particles have no volume and are not attracted to each other

  - REAL GAS – does not follow gas laws exactly, it deviates at low temperatures and high pressures
    - Condenses to liquid or sometimes solid when cooled or under pressure
    - Particles are attracted to each other and have volume
    - Behaves like an ideal gas at higher temperatures and lower pressures

Using the Ideal Gas Law

Example One: What mass of neon gas should be introduced into an evacuated 0.88L tube to produce a pressure of 90 kPa at 30°C?

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(90\text{kPa})(0.88\text{L})}{(8.314\ \text{L}\bullet\text{kPa/mol}\bullet\text{K})(303\text{K})}$$

$$n = 0.0314\ \text{mol} \times \frac{20.18\ \text{g}}{1\ \text{mol}} = 0.63\ \text{g}$$

Using the Ideal Gas Law – Practice

3. Predict the volume occupied by 0.78 g of hydrogen at 22°C and 125 kPa

$$n\ H_2 = 0.78\ g \times \frac{1\ mol}{2.02g} = 0.386\ mol$$

$$PV = nRT \rightarrow \quad V = \frac{nRT}{P}$$

$$V = \frac{(0.386\ mol)(8.314 L\bullet kPa/mol\bullet K)(295K)}{125\ kPa}$$

$V = 7.573$ L

V = 7.6 L

SOLUTIONS

Dissociation vs. Ionization

- **What is the difference between dissociation and ionization?**
- Both produce (aq) ions...
- Dissociation, however, is the separation of ions that already exist before dissolving in water

$$M^+X^-_{(s)} \rightarrow M^+_{(aq)} + X^-_{(aq)}$$

- Ionization involves the production of new ions, specifically hydrogen ions

$$HX^0_{(aq)} \rightarrow H^+_{(aq)} + X^-_{(aq)}$$

SUMMARY

Substance	Process	General Equation
Molecular	Disperse as individual, neutral molecules	$XY\ (s/l/g) \rightarrow XY\ (aq)$
Ionic	Dissociate into individual ions	$MX\ (s) \rightarrow M^{+}{}_{(aq)} + X^{-}{}_{(aq)}$
Base (ionic hydroxide)	Dissociate into positive ions and hydroxide ions	$MOH_{(s)} \rightarrow M^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$
Acid	Ionize to form hydrogen ions and anions	$HX_{(s/l/g)} \rightarrow H^{+}{}_{(aq)} + X^{-}{}_{(aq)}$

Reference pg. 201

1) **Percent Concentration (%)**

a) Percentage by volume:
(volume of solute(mL) in 100mL)

$$c = \frac{V_{solute}}{V_{solution}} \times 100\%$$

b) Percentage weight by volume:
(mass of solute(g) in 100mL)

$$c = \frac{m_{solute}}{V_{solution}} \times 100\%$$

c) Percentage weight by weight:

$$c = \frac{m_{solute}}{m_{solution}} \times 100\%$$

1) **Parts per million (ppm)** – the <u>mass in mg of solute</u> in every L of solution

- This unit gives more reasonable numbers to very dilute concentrations (one ppm is only a drop of solute in a full bathtub)
- *<u>For aqueous solutions</u>*: 1 ppm = 1 g/10^6 mL = 1 mg/L = 1 mg/kg

$$c = \frac{m_{\text{solute (g)}}}{m_{\text{solution (g)}}} \times 10^6 \qquad \text{or} \qquad c = \frac{m_{\text{solute (mg)}}}{m_{\text{solution (kg)}}}$$

Bicarbonates: 320 PPM
Calcium Ion: 150 PPM
Fluoride: 0.12 PPM
Magnesium: 4.2 PPM
Potassium: 1.2 PPM
Sulfates: 43 PPM
Sodium: 11 PPM

- **Amount Concentration (aka Molar Concentration) (mol/L)** – the amount of moles of solute per litre of solution; this is the most common form of concentration used in chemistry

- $c = \dfrac{n_{\text{solute}}\ (\text{mol})}{V_{\text{solution (L)}}}$

Guided Examples:

1) Find the amount concentration of aluminum and sulphate ions in a 0.30 mol/l solution of aluminum sulphate

- $Al_2(SO_4)_{3\,(aq)} \rightarrow 2\,Al^{3+}{}_{(aq)} + 3\,SO_4{}^{2-}{}_{(aq)}$
- c = 0.30 mol/L

- $[Al^{3+}{}_{(aq)}] = 0.30\text{ mol/L} \times \frac{(2)}{1} = 0.60\text{ mol/L}$

- $[SO_4{}^{2-}{}_{(aq)}] = 0.30\text{ mol/L} \times \frac{(3)}{1} = 0.90\text{ mol/L}$

1. A vinegar solution contains 4.30 g of acetic acid, $CH_3COOH(aq)$, in 200 g of solution. The percent by mass concentration is:

   a) 20.3% m/m
   b) 14.8% m/m
   c) 2.15% m/m
   d) 0.0493% m/m

2. The concentration of lead ions, $Pb^{2+}(aq)$, in bottled water should not exceed 0.20 ppm. The maximum mass of lead ions that can be dissolved in 500 mL of water is:

   a) 0.10 mg
   b) 0.38 mg
   c) 0.60 mg
   d) 0.72 mg

3. Calculate the volume of a 0.344 mol/L HCl solution that contains 0.12 mol of solute. The volume would be:

   a) 0.04 L
   b) 0.34 L
   c) 0.35 L
   d) 2.87 L

4. What is the molar concentration if you dissolve 6.00 g of NaCl in 200 mL of water? (Remember M = mol/L)

   a) 0.0200 M
   b) 0.513 M
   c) 1.20 M
   d) 5.84 M

5. If the concentration of a solution of magnesium iodide is 0.15 mol/L solution, then the concentration of the iodide ions is:

   a) 0.15 mol/L
   b) 0.24 mol/L
   c) 0.30 mol/L
   d) 0.60 mol/L

6. The mass of sodium sulfate (Na_2SO_4) needed to make 250 mL of a 0.129 mol/L solution is:

   a) 142 g
   b) 10.9 g
   c) 4.58 g
   d) 3.23 g

Determining the mass of pure solid for a Standard Solution

- Use conversion factors to determine the values for both the amount in moles and the mass of solid required.
- Because you are working with one substance, you do not need a balanced equation (No need for a mol ratio)
- Volume of the solution and its molar concentration are needed.
- Example: To prepare 250.0mL of 0.100 mol/L solution of sodium carbonate, the mass needed is:

$$0.2500\ \text{L} \times \frac{0.100\ \text{mol}}{1\ \text{L}} \times \frac{105.99\ \text{g}}{1\ \text{mol}} = 2.65\ \text{g}$$

Determining the volume of Stock Solution for a Standard Solution

- We know that the number of moles does not change when diluting a solution, but the concentration and volume will. So we will use the dilution formula:

$$\mathbf{C_1V_1 = C_2V_2}$$

- To use the dilution formula, you must make sure the units are consistent for both the c and v (i.e. both in mL or both in L)
- Example: How would you prepare 100 mL of 0.40 mol/L $MgSO_{4(aq)}$ from a solution of 2.0 mol/l $MgSO_{4(aq)}$

$$C_1V_1 = C_2V_2$$

$$(2.0\text{mol/L})\ (V_1) = (0.40\text{mol/L})\ (100\text{mL})$$

$$V_1 = 20\text{ mL}$$

4. What volume of 75 mmol/L solution can be prepared from 10 g of sodium carbonate?

5. Determine the volume of concentrated hydrochloric acid required to prepare 10.0 L of a 0.200 mol/L solution.

ACIDS AND BASES

Empirical Definitions

Acid – a substance which dissolves in water to produce a solution that:

- Tastes sour
- Turns blue litmus red
- Conducts electricity
- Reacts with active metals to produce $H_2(g)$
- Neutralizes Bases

Base – a substance which dissolves in water to produce a solution that:

- Tastes bitter; feels slippery
- Turns red litmus blue
- Conducts electricity
- Neutralizes acids

Theoretical Definitions

a) <u>Arrhenius:</u>

Acid – a substance that forms an acidic solution by dissolving in water to produce free hydrogen ions (**$H^+_{(aq)}$**) in solution

- Example: $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

Base – a substance that forms a basic solution by dissolving in water to produce free hydroxide ions (**$OH^-_{(aq)}$**) in solution

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

Theoretical Definitions

b) Modified Definition:

Acid – a species that forms an acidic solution by reacting with water to produce hydronium ions (**$H_3O^+_{(aq)}$**)

- Example: $HCl_{(aq)} + H_2O_{(aq)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

Base – a species that forms a basic solution by reacting with water to produce hydroxide ions (**$OH^-_{(aq)}$**)

- Example: $NH_{3\,(aq)} + H_2O_{(aq)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$

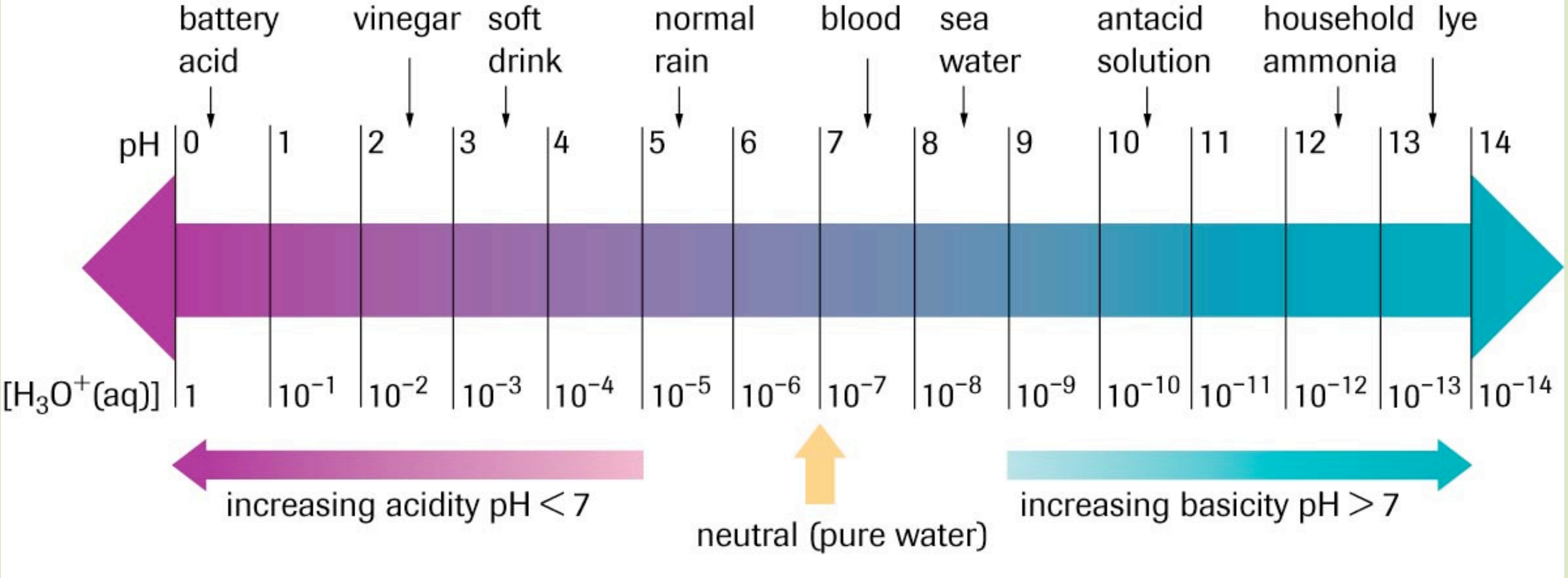
The **hydronium ion** (hydrated proton) – was discovered by **Paul Giguère** at the Université Laval in 1957.

Practice – Naming Acids

	IUPAC	Traditional
– $HClO_{4(aq)}$	aqueous hydrogen perchlorate	
– $HClO_{(aq)}$	aqueous hydrogen hypochlorite	
– $HNO_{2(aq)}$		nitrous acid
– $HNO_{3(aq)}$		nitric acid
– $H_2SO_{4(aq)}$	aqueous hydrogen sulfate	sulfuric acid
– $H_2SO_{3(aq)}$	aqueous hydrogen sulfite	sulfurous acid
– $H_3PO_{4(aq)}$	aqueous hydrogen phosphate	phosphoric acid

pH – power of hydrogen

$$[H_3O^{+}{}_{(aq)}] = 10^{-pH}$$



The pH scale is used to communicate a broad range of hydronium ion concentrations. Most common acids and bases have pH values between 0 and 14

Summary

$\boxed{pH = -\log [H_3O^+_{(aq)}]}$

$\boxed{[H_3O^+_{(aq)}] = 10^{-pH}}$

$\boxed{pOH = -\log [OH^-_{(aq)}]}$

$\boxed{[OH^-_{(aq)}] = 10^{-pOH}}$

- The number of digits **following the decimal point** in a pH or pOH value is equal to the number of significant digits in the corresponding hydronium or hydroxide concentration.

- For both pH and pOH, an **inverse** relationship exist between the ion concentration and the pH or pOH. The greater the hydronium ion concentration, the lower the pH is.

Practice

- Pg. 242 #4-7 (pH)
- Pg. 243 #9-11 (pOH)

Acid – Base Indicators

- Other acid-base indicators:
  - Because the chemical structure of each indicator is different, the pH at which the indicator changes from the $HIn_{(aq)}$ form to the $In^{-}_{(aq)}$ form is different for each indicator. *(See inside back cover of textbook)*

Common name	Color of HIn(aq)	pH range of colour change	Color of In^{-} (aq)
Bromothymol blue	Yellow	6.0-7.6	Blue
Phenolphthalein	Colourless	8.2-10.0	Pink

Practice

- Lab Exercise 6.B pg. 247
  - Complete the Analysis Portion of the lab below

Solution	**After addition to samples of the solution …**	**pH range implied by each indicator**	**pH of solution**
A	● methyl violet was blue ● methyl orange was yellow ● methyl red was red ● phenolphthalein was colourless		
B	● indigo carmine was blue ● phenol red was yellow ● bromocresol green was blue ● methyl red was yellow		
C	● phenolphthalein was colourless ● thymol blue was yellow ● bromocresol green was yellow ● methyl orange was orange		

More Practice....

- $HI_{(aq)}$ – explain acidic properties
  - $HI_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + I^-_{(aq)}$
- $NaCH_3COO_{(aq)}$ – explain basic properties
  - $NaCH_3COO_{(aq)} \rightarrow Na^+_{(aq)} + CH_3COO^-_{(aq)}$ Simple Dissociation
  - $CH_3COO^-_{(aq)} + H_2O_{(l)} \rightarrow CH_3COOH_{(aq)} + OH^-_{(aq)}$
- $HOCl_{(aq)}$ – explain acidic properties
  - $HOCl + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + OCl^-_{(aq)}$

Try Three More…

- $H_3PO_{4(aq)}$ – explain acidic properties
  - $H_3PO_{4(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + H_2PO_4^-{}_{(aq)}$

- $Na_2SO_{4\,(aq)}$ – explain basic properties
  - $Na_2SO_{4\,(aq)} \rightarrow 2\,Na^+_{(aq)} + SO_4^{2-}{}_{(aq)}$ Simple Dissociation
  - $SO_4^{2-}{}_{(aq)} + H_2O_{(l)} \rightarrow HSO_4^-{}_{(aq)} + OH^-_{(aq)}$

- $Sr(OH)_{2(aq)}$ – explain basic properties
  - $Sr(OH)_{2(aq)} \rightarrow Sr^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

Summary

- Acids are substances that react with water to produce hydronium ions
- Most bases are substances that react with water to produce hydroxide ions
- Neutralization can be explained as the reaction between hydronium ions and hydroxide ions to produce water.
- Try pg. 253 #4,5

The Difference: Using the Modified Arrhenius Theory

- **<u>Strong Acids</u>: have high conductivity, high rate of reaction w/ metals and carbonates and a relatively low pH**

- **These empirical properties suggest many ions are present (lots of H_3O^+ ions present); which is consistent with the idea that strong acids react completely (>99%) with water to form hydronium ions**

$$HCl_{(aq)} + H_2O_{(l)} \xrightarrow{>99\%} H_3O^+_{(aq)} + Cl^-_{(aq)}$$

The Difference: Using the Modified Arrhenius Theory

- **<u>Weak Acids</u>: have low conductivity, a lower rate of reaction w/ active metals and carbonates and a relatively high pH**

- **These empirical properties suggest fewer hydronium ions are present**

- **Based on this evidence, a weak acid reacts incompletely (<50%) with water to form relatively few hydronium ions**

$$CH_3COOH_{(aq)} + H_2O_{(l)} \xrightarrow{<50\%} H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

Strong and Weak Bases

- Strong Bases – all soluble ionic hydroxides that dissociate completely (>99%) to release hydroxide ions

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

- Weak Bases – an ionic or molecular substance that reacts partially (<50%) with water to produce relatively few hydroxide ions

<50%

$$NH_{3\,(aq)} + H_2O_{(aq)} \rightarrow OH^{-}_{(aq)} + NH_4^{+}{}_{(aq)}$$

Summary

	Strong Acids	**Weak Acids**	**Strong Bases**	**Weak Bases**
Empirical properties (need same concentration & temperature)	Very low pH	Med to low pH	Very high pH	Med to high pH
	High conductivity	Low conductivity	High conductivity	Low conductivity*
	Fast reaction rate	Slow reaction rate	Fast reaction rate	Slow reaction rate
Modified Arrhenius Theory	Completely react with water to form $H_3O^+_{(aq)}$ ions	Partially react with water to form $H_3O^+_{(aq)}$ ions	Completely react with water to form $OH^-_{(aq)}$ ions	Partially react with water to form $OH^-_{(aq)}$ ions

* Applies only to weak bases that are molecular

Polyprotic Acids

- With all of these possible reactions you might think that phosphoric acid is very acidic, but only the first reaction is significant.

- Also, the first reaction is only 50% complete, so it is actually a weak acid and it has a pH noticeable higher than a strong acid at the same concentration.

Acid	Concentration	pH
$HCl_{(aq)}$	0.1 mol/L	1.0
$H_3PO_{4(aq)}$	0.1 mol/L	1.7

- Based on this evidence:
  - **In general, polyprotic acids are weak acids whose reaction with water decreases with each successive step.**

NOTE: $H_2SO_{4(aq)}$ is a notable exception to this rule. It is a strong acid because it's first reaction with water is essentially complete. However the second reaction is much less than 50% complete.

2. What happens to the pH of a solution when:

   (a) the hydronium ion concentration increases?

   (b) the hydronium ion concentration decreases?

   (c) a strong acid is diluted?

3. Some hydrochloric acid is completely neutralized with a sodium hydroxide solution. Describe the pHs of the reactants and final solution.

4. A vinegar solution has a pH of 2.82. Calculate the hydronium ion concentration of this solution.

STOICHIOMETRY

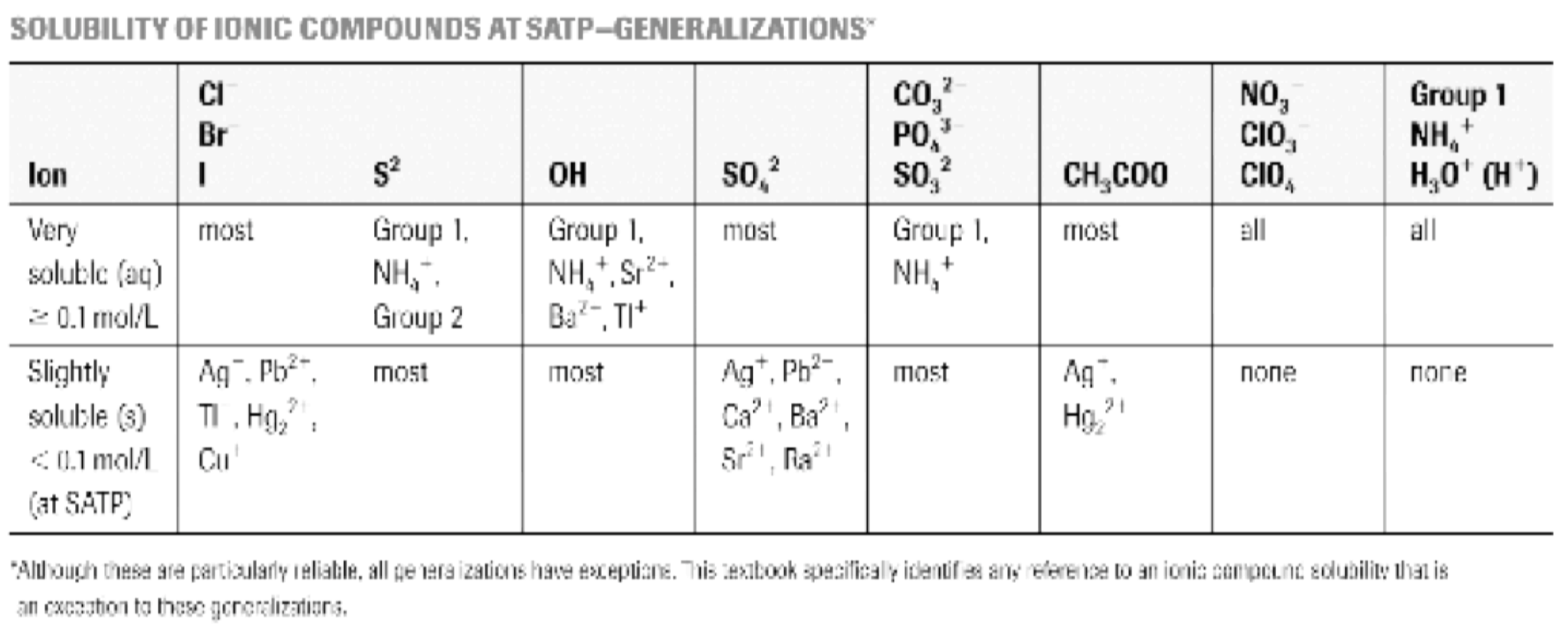
What do you remember?

In Science 10 you learned about five reaction types, can you match them up

- Composition (Formation)
- Decomposition
- Combustion
- Single Replacement
- Double Replacement

- $CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$
- $Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$
- $Cu_{(s)} + AgNO_{3(aq)} \rightarrow Ag_{(s)} + Cu(NO_3)_{2(g)}$
- $CaCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow CaCO_{3(s)} + NaCl_{(aq)}$
- $H_2O_{(l)} \rightarrow O_{2(g)} + H_{2(g)}$

Using the solubility table:



SOLUBILITY OF IONIC COMPOUNDS AT SATP—GENERALIZATIONS*

Ion	**Cl^-** **Br^-** **I^-**	**S^{2-}**	**OH^-**	**SO_4^{2-}**	**CO_3^{2-}** **PO_4^{3-}** **SO_3^{2-}**	**CH_3COO^-**	**NO_3^-** **ClO_3^-** **ClO_4^-**	**Group 1** **NH_4^+** **H_3O^+ (H^+)**
Very soluble (aq) ≥ 0.1 mol/L	most	Group 1, NH_4^+, Group 2	Group 1, NH_4^+, Sr^{2+}, Ba^{2+}, Tl^+	most	Group 1, NH_4^+	most	all	all
Slightly soluble (s) < 0.1 mol/L (at SATP)	Ag^+, Pb^{2+}, Tl^+, Hg_2^{2+}, Cu^+	most	most	Ag^+, Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}	most	Ag^+, Hg_2^{2+}	none	none

*Although these are particularly reliable, all generalizations have exceptions. This textbook specifically identifies any reference to an ionic compound solubility that is an exception to these generalizations.

Practice

When cancelling spectator ions, they must be identical in every way: chemical amount, form (atom, ion, molecule) and state of matter

- Write the net ionic equation for the reaction of aqueous barium chloride and aqueous sodium sulfate. (Refer to the solubility table)
- 1) $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$
- 2) $Ba^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2Na^{+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)} + 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)}$

  (Complete ionic equation)
- 3) $Ba^{2+}_{(aq)} + \cancel{2Cl^{-}_{(aq)}} + \cancel{2Na^{+}_{(aq)}} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)} + \cancel{2Na^{+}_{(aq)}} + \cancel{2Cl^{-}_{(aq)}}$
- 4) $Ba^{2+}_{(aq))} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$ **(Net ionic equation)**
- Ions that are present but do not take part in (change during) a reaction are called spectator ions (like spectators at a sports game: they are present but do not take part in the game)

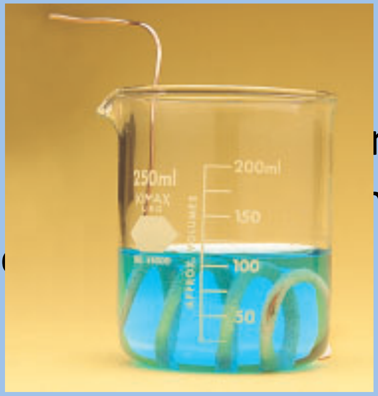
Limiting and Excess Reagents

- When no further changes appear to be occurring, we assume that all of the $AgNO_{3(aq)}$ that was initially present has now been completely reacted.

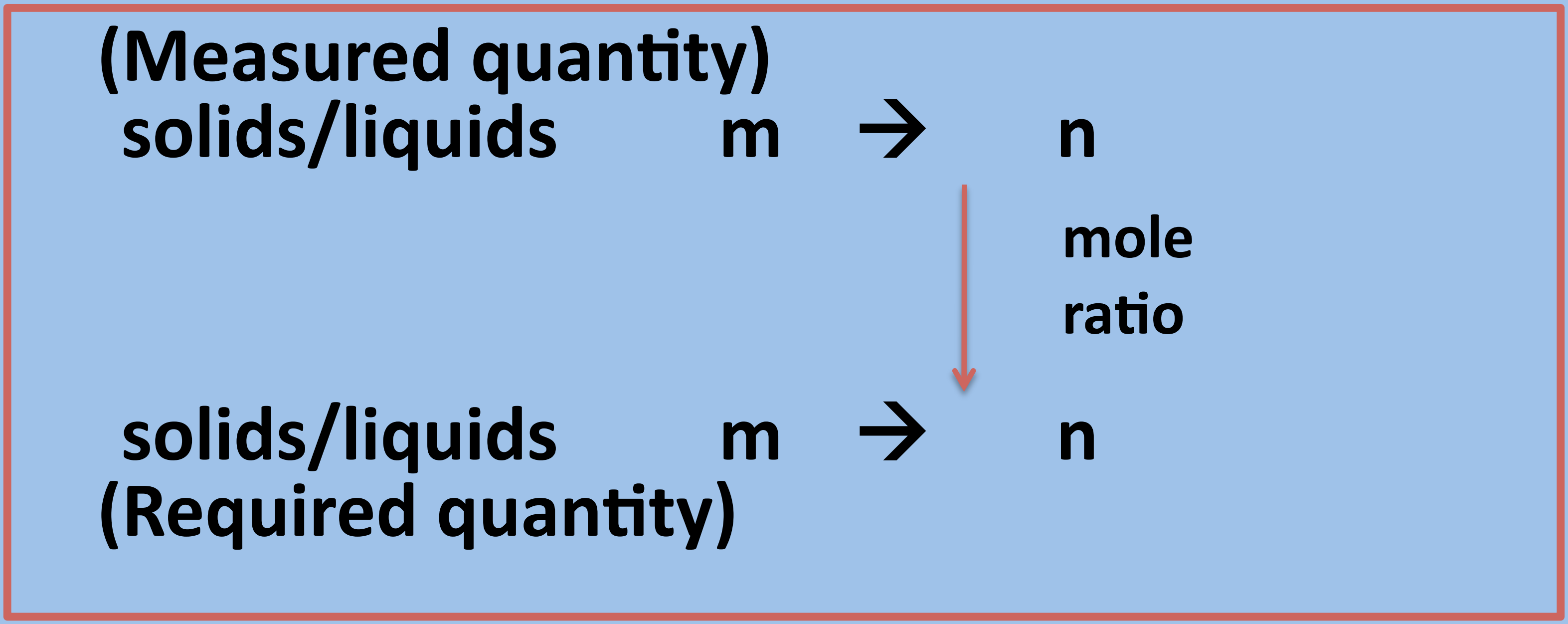
- A **limiting reagent** is the reactant whose entities are completely consumed in a reaction, meaning the reaction stops.
    - In order to make sure this happens, more of the other reactant must be present than is required

- An **excess reagent** is the reactant whose entities are present in surplus amounts, so that some remain after the reaction ends..

- <u>In our reaction</u>: much more copper was used than nced by the unreacted copper) so we assume the hen no more silver ions were left, so silver nitrat reagent.



Stoichiometry Calculations



(Measured quantity)

solids/liquids m → n

mole ratio

solids/liquids m → n

(Required quantity)

Practice #3 (Mass Stoichiometry)

What mass of iron (III) oxide is required to produce 100.0 g of iron?

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

m = ?	m = 100.0g
M = 159.70g/mol	M = 55.85 g/mol

$$m\ Fe_2O_{3(s)}:\ 100.0\ g \times \frac{1\ mol}{55.85\ g} \times \frac{1\ mol}{2\ mol} \times \frac{159.70\ g}{1\ mol} = 143.0\ g\ Fe_2O_3$$

Percent Yield for Reactions

- We can use stoichiometry to test experimental designs, technological skills, purity of chemicals, etc. We evaluate these by calculating a **percent yield**.
  - This is the ratio of the *actual (experimental) quantity* of product obtained and the *theoretical (predicted) quantity* of product obtained from a stoichiometry calculation
  - **Percent yield = $\frac{\text{actual yield}}{\text{predicted yield}} \times 100$**
  - Some forms of experimental uncertainties:
    - All measurements (limitations of equipment)
    - Purity of chemical used (80-99.9% purity)
    - Washing a precipitate (some mass is lost through filter paper)
    - Estimation of reaction completion (qualitative judgements i.e. color)

Percent Yield Example #1

- Example: In a chemical analysis, 3.00 g of silver nitrate in solution was reacted with excess sodium chromate to produced 2.81 g of filtered, dried precipitate.

- **Predicted value**: $2AgNO_{3(aq)} + Na_2CrO_{7(aq)} \rightarrow Ag_2CrO_{7(s)} + 2NaNO_{3(aq)}$

  m = 3.00 g　　　　m =

  M = 169.88g/mol　　　　M = 331.74 g/mol

$$3.00\text{g} \times \frac{1\text{ mol}}{169.88\text{g}} \times \frac{1\text{ mol}}{2\text{ mol}} \times \frac{331.74\text{ g}}{1\text{ mol}} = 2.93\text{ g}$$

- **Percent yield** = $\frac{\text{actual yield}}{\text{predicted yield}} \times 100\% = \frac{2.81\text{g}}{2.93\text{g}} \times 100\% = 95.9\%$

Gas Stoichiometry

- If 300g of propane burns in a gas barbecue, what volume of oxygen measured at SATP is required for the reaction?
- Remember: 24.8L/mol for SATP

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

m = 300g

44.11g/mol

V = ?

24.8L/mol

$$300\ g \times \frac{1\ mol}{44.11\ g} \times \frac{5\ mol}{1\ mol} \times \frac{24.8\ L}{1\ mol} = 843\ L\ O_{2(g)}$$

****Remember – molar volume is the conversion factor for gases just like molar mass is the conversion factor in gravimetric stoichiometry**

Gas Stoichiometry

- What volume of ammonia at 450kPa and 80°C can be obtained from the complete reaction of 7.5kg of hydrogen with nitrogen?

$$2N_{2(g)} \quad + \quad 3H_{2(g)} \quad \rightarrow \quad 2NH_{3(g)}$$

m = 7500g
M = 2.02 g/mol

V = ?
P = 450kPA
T = 353.13K

$$7500\text{ g} \times \frac{1\text{ mol}}{2.02\text{ g}} \times \frac{2}{3} = 2475.2475\text{ mol } NH_{3(g)}$$

$$PV = nRT \rightarrow V = \frac{nRT}{P} = \frac{(2475.2475\text{ mol})(8.314^{\text{kpa}\bullet\text{L}}/_{\text{mol}\bullet\text{K}})(353.15\text{K})}{(450\text{kPa})}$$

$$= 16150.10\text{L} \rightarrow \mathbf{1.6 \times 10^{4}\ L\ of\ NH_{3(g)}}$$

Solution Stoichiometry

- In an experiment, a 10.00 mL sample of sulfuric acid solution reacts completely with 15.9 mL of 0.150 mol/L potassium hydroxide. Calculate the amount concentration of the sulfuric acid.

- $H_2SO_{4(aq)} + 2KOH_{(aq)} \rightarrow 2H_2O_{(l)} + K_2SO_{4(aq)}$

  V = 10.00mL   V = 15.9 mL

  c = ?   0.150 mol/L

$$15.9\text{mL} \times \frac{0.150\text{ mol}}{1\text{ L}} \times \frac{1\text{mol}}{2\text{ mol}} \times \frac{1}{10.0\text{ mL}} = 0.119\text{ mol/L}$$

Chemical Analysis

Colorimetry

Ion	**Solution colour**
Group 1, 2, 17	Colourless
$Cr^{2+}_{(aq)}$	Blue
$Cr^{3+}_{(aq)}$	Green
$Co^{2+}_{(aq)}$	Pink
$Cu^{+}_{(aq)}$	Green
$Cu^{2+}_{(aq)}$	Blue
$Fe^{2+}_{(aq)}$	Pale-green
$Fe^{3+}_{(aq)}$	Yellow-brown
$Mn^{2+}_{(aq)}$	Pale pink
$Ni^{2+}_{(aq)}$	Green
$CrO_4^{2-}{}_{(aq)}$	Yellow
$Cr_2O_7^{2-}{}_{(aq)}$	Orange
$MnO_4^{-}{}_{(aq)}$	Purple

Which solution is which?

potassium dichromate	6
sodium chloride	2
sodium chromate	4
potassium permanganate	1
nickel (II) nitrate	3
copper (II) sulfate	5

1) You want to test the stoichiometric method using the reaction of 2.00 g of copper(II) sulfate in solution with an excess of sodium hydroxide in solution. What would be a reasonable mass of sodium hydroxide to use?
   - *To answer this question, you need to calculate the minimum mass required and then add 10%.*

$$CuSO_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + Na_2SO_{4(aq)}$$

2.00g *m = ?*

159.62g/mol *40.00g/mol*

$$2.00\text{ g} \times \frac{1\text{ mol}}{159.62\text{ g}} \times \frac{2}{1} \times \frac{40.0\text{g}}{1\text{ mol}} = 1.00\text{ g} \qquad 10\% = 0.10\text{g}$$

1.00 g x 1.10%= **1.10g**

Practice pg. 321 #2

2) If 10.0g of copper is placed in solution of 20.0g of silver nitrate, which reagent will be the limiting reagent?

- All reactants must be converted to moles, then using the mole ratio, determine which reactant will run out first.

$$Cu_{(s)} \quad + \quad 2\ AgNO_{3(aq)} \rightarrow 2\ Ag_{(s)} \quad + \quad Cu(NO_3)_{2(aq)}$$

10.0g 20.0g

given: 0.157 mol 0.0589 mol

$$n_{Cu(s)}\text{: } 10.0\text{g} \times \frac{1\text{ mol}}{63.55\text{ g}} = 0.157\text{ mol} \times \frac{2}{1} = 0.315\text{ mol}$$

That much silver nitrate is not available so copper is not the limiting reagent

$$n_{AgNO3}\text{: } 20.0\text{g} \times \frac{1\text{ mol}}{169.88\text{ g}} = 0.118\text{ mol} \times \frac{1}{2} = 0.0589\text{ mol}$$

More copper than that is available so silver nitrate is the limiting reagent

You must test one of the values using the mole ratio. Assume one chemical is completely used up and see if enough of the second chemical is present.

Limiting and Excess Reagents Summary

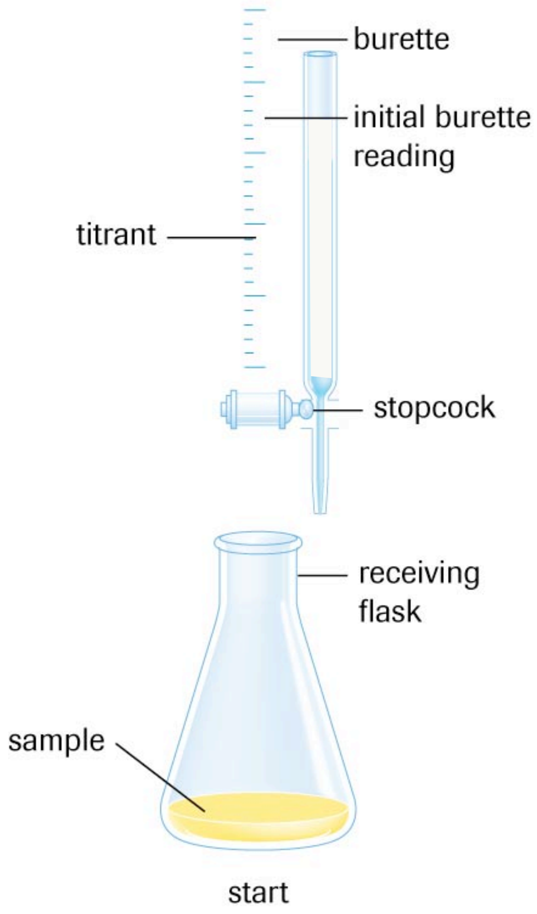
- Identify the limiting reagent by choosing either reagent amount, and use the mole ratio to compare the required amount with the amount actually present.

- The quantity in excess is the difference between the amount of excess reagent present and the amount required for complete reaction.

- A reasonable reagent excess to use to ensure complete reaction is 10%.

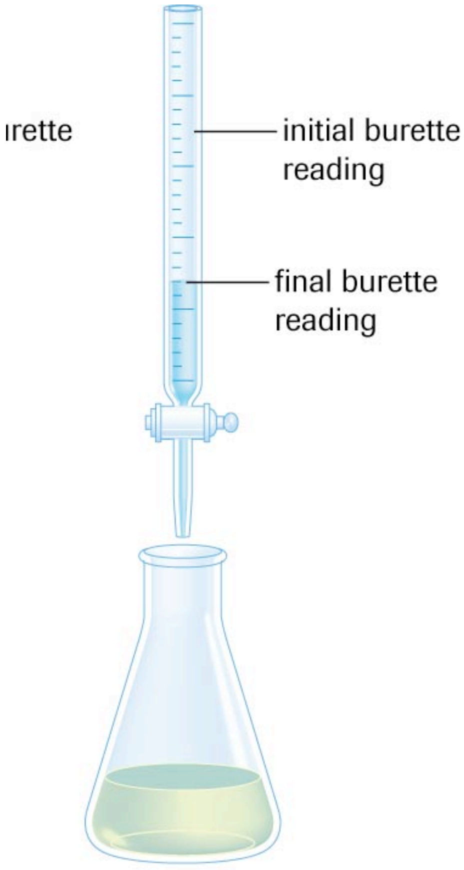
Chemical Analysis by Titration

- **Titration** – is a common experimental design used to determine the amount concentration of substances in solution.
  - The solution of known concentration may be either the titrant or the sample; it makes no difference to the analysis
- Titration breakdown:
  - Carefully adding a solution (**titrant)** from a *burette* into a measured, fixed volume of another solution (**sample**) in an *Erlenmeyer flask* until the reaction is judged to be complete



Chemical Analysis by Titration

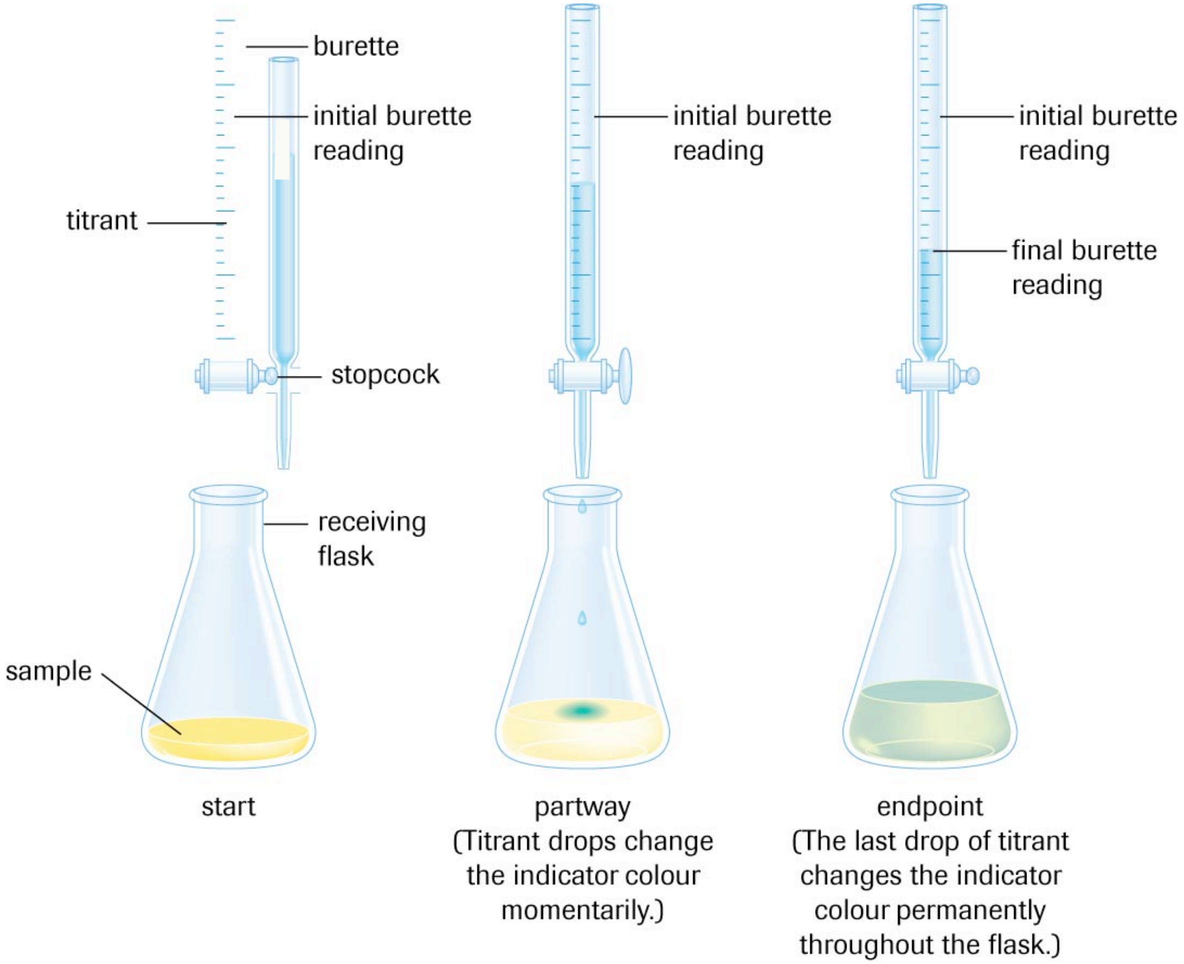
- **Burette** – precisely marked glass cylinder with a stopcock at one end. Allows precise, accurate measurement and control of the volume of reacting solution.
- When doing a titration, there will be a point at which the reaction is complete; *when chemically equivalent amounts of reactants have combined*. This is called the <u>equivalence point</u>:
  - **Equivalence point** – the point during a titration at which the exact theoretical chemical amount of titrant has been added to the sample. (QUANTITATIVE)
- To measure this equivalence point experimentally, we look for a sudden change in an observable property, such as color, pH, or conductivity. This is called the **endpoint**. (QUALITATIVE)



endpoint
(The last drop of titrant changes the indicator colour permanently throughout the flask.)

Chemical Analysis by Titration

- An <u>initial reading</u> of the burette is made before any titrant is added to the sample.
- Then the titrant is added until the reaction is complete; when a final drop of titrant *permanently* changes the colour of the sample.
- The <u>final burette reading</u> is then taken.
- The <u>difference</u> between the readings is the volume of titrant added.



Near the endpoint, continuous gentle swirling of the solution is important

Sample Problem

- Determine the concentration of hydrochloric acid in a commercial solution.
- A 1.59g mass of sodium carbonate, $Na_2CO_{3(s)}$, was dissolved to make 100.0mL of solution. Samples (10.00mL) of this standard solution were then taken and titrated with the hydrochloric acid solution.

- The titration evidence collected is below. Methyl orange indicator was used.

Trial	**1**	**2**	**3**	**4**
Final burette reading (mL)	13.3	26.0	38.8	13.4
Initial burette reading (mL)	0.2	13.3	26.0	0.6
Volume of $HCl_{(aq)}$ added	13.1	12.7	12.8	12.8
Indicator colour	Red	Orange	Orange	Orange

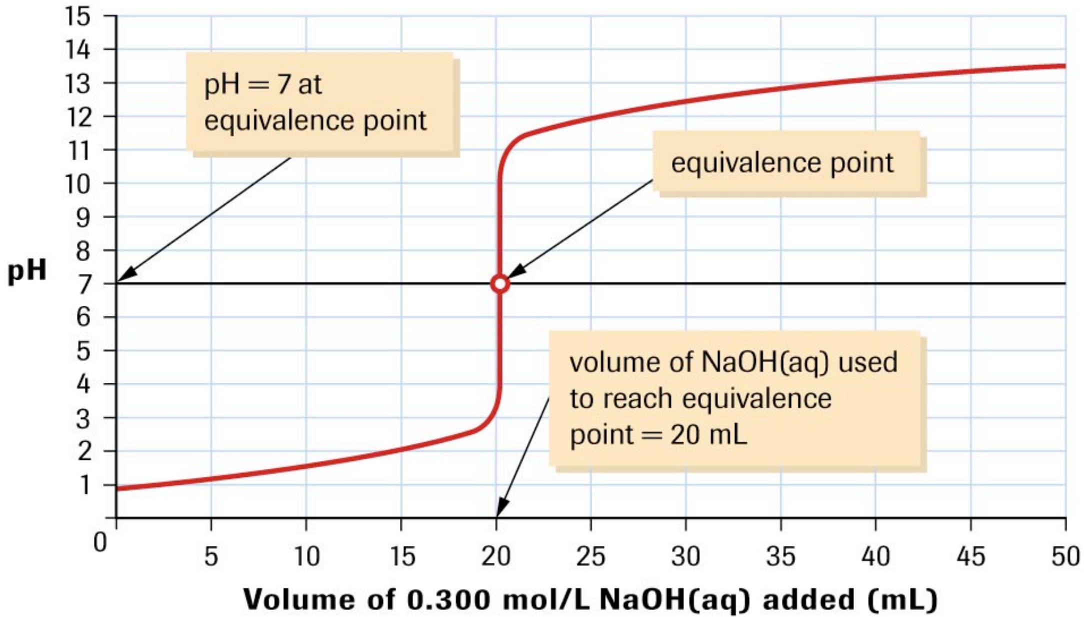
TIP: In titration analysis, the first trial is typically done very quickly. It is just for practice, to learn what the endpoint looks like and to learn the approximate volume of titrant needed to get to the endpoint. Greater care is taken with subsequent trials

- The initial addition of the titrant (in the burette) to the acid does not produce large changes. This relatively flat region of the pH curve is where a <u>buffering action</u> occurs.

- As the titration proceeds, and base is added, some of the acid is reacted with the added base, but anywhere before the equivalence point some <u>excess acid</u> will remain, so the pH stays relatively low.

- Very near the equivalence point, a small excess of acid becomes a small excess of base with the addition of a few more drops, so the pH <u>abruptly</u> changes.

- The equivalence point is the <u>centre</u> of this change, where the curve is the most vertical.



Titration Curve for Titration of 20 mL of 0.300 mol/L HCl(aq) with 0.300 mol/L NaOH(aq)

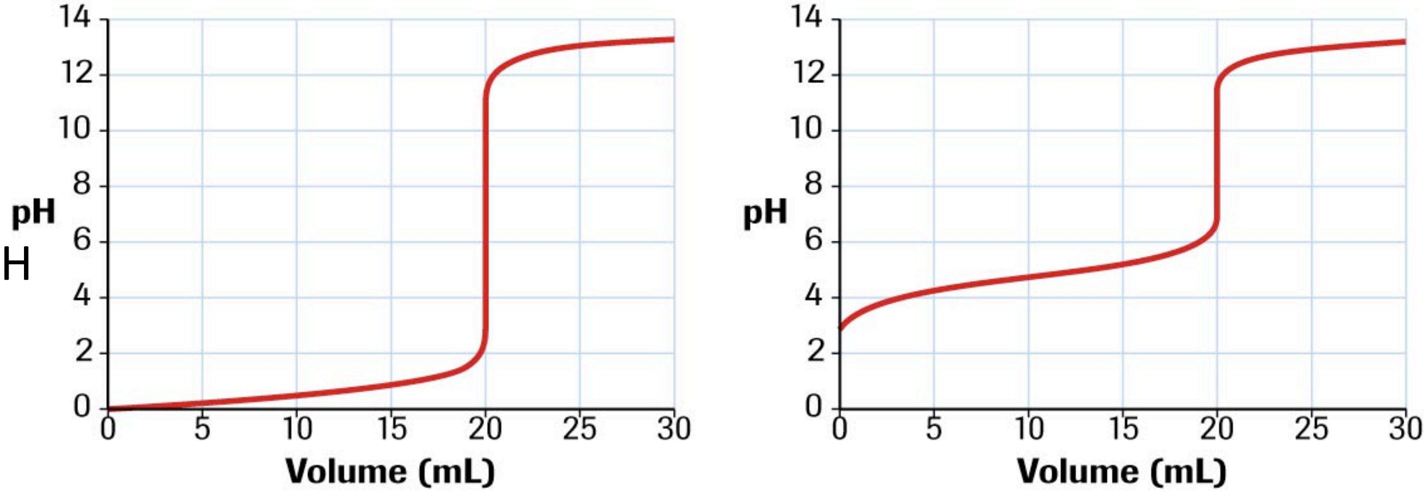
General Rule

<u>Strong Acid to Weak Base</u>: pH at equivalence point is always lower than 7

<u>Strong Base to Weak Acid</u>: pH at equivalence point is always higher than 7

Titrations Involving Monoprotic Acids and Bases of Equal Concentration

(a) Strong Base Added to Strong Acid (b) Strong Base Added to Weak Acid



(c) Strong Acid Added to Weak Base (d) Strong Acid Added to Strong Base