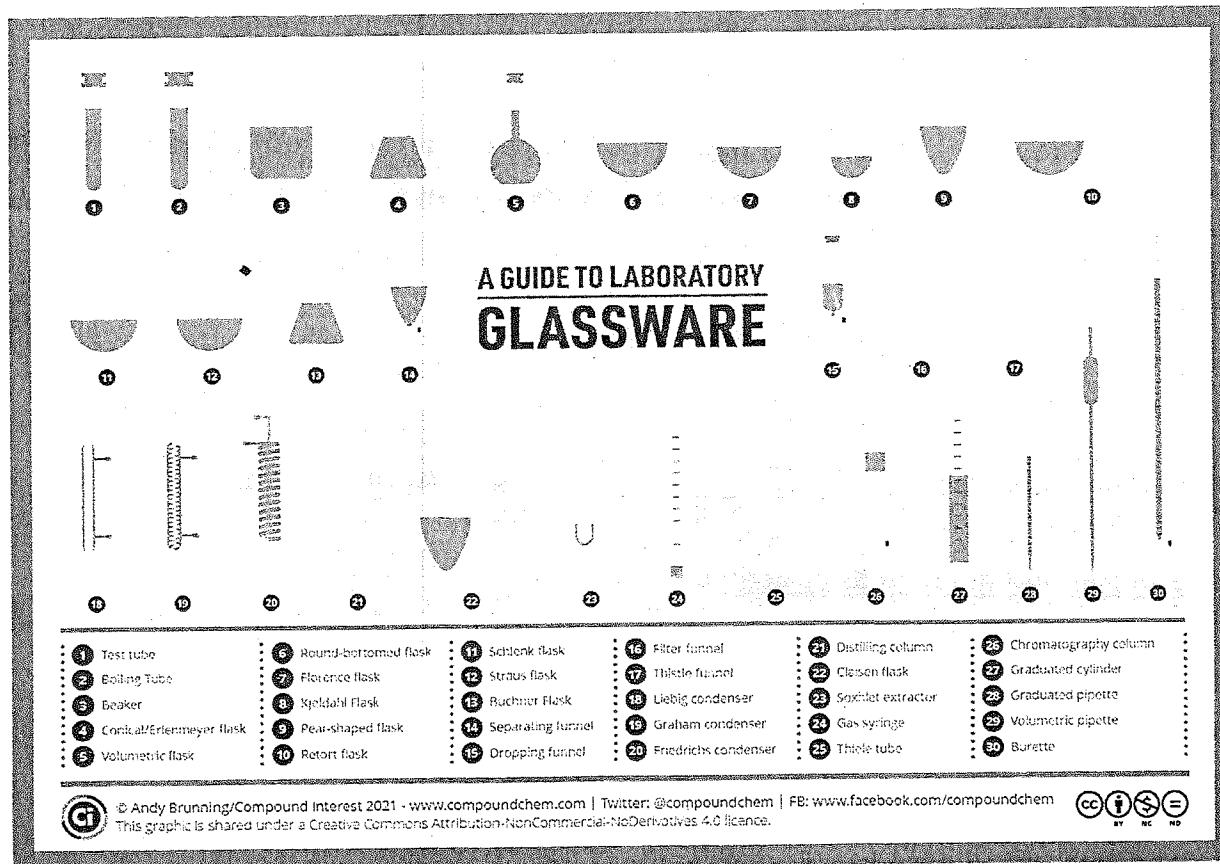


Grade 12 University Level Chemistry

INTRODUCTION UNIT

Lesson	Learning Goals
Math Skills	<input checked="" type="checkbox"/> Convert numbers from standard form to scientific notation and vice versa <input checked="" type="checkbox"/> Identify the correct significant digits when performing math equations <input checked="" type="checkbox"/> Demonstrate units using dimensional analysis
Stoichiometry	<input type="checkbox"/> Identify a mole ratio between reactants and products <input type="checkbox"/> Use dimensional analysis to convert from known mass to an unknown <input type="checkbox"/> Use dimensional analysis to convert from known concentration and volume to an unknown
Limiting & Excess Reagents	<input type="checkbox"/> Identify limiting and excess reagents <input type="checkbox"/> Use limiting reagents to determine values for unknown chemical species <input type="checkbox"/> Calculate moles remaining
Titrations	<input type="checkbox"/> Define acids, bases and titrations <input type="checkbox"/> Calculate the pH and pOH of strong acids and bases <input type="checkbox"/> Use mole ratios to determine concentrations and pH



SCH4U Math Skills

Significant Digits: Precision is very important in chemistry. Each digit obtained as a result of a measurement is called a significant figure. For example the measurement 6.77 cm has three significant figures, while 45 m has only two. Every number that you have seen has a certain number of significant digits. There are 5 rules to identifying sig digs:

Identifying Significant Digits

Rule #1: "Non-zero numbers are significant"

How many sig figs are there in the numbers below?

3.1428 5

3.14 3

469 3

Rule #2: "Sandwiched zeros are significant!" All zeros between non-zero numbers are significant.

How many sig figs are there in the numbers below?

7.053×10^3 4

70501 5

302 3

Rule #3: "Zeros on the left are NOT significant" Zeros to the left of the first significant figure are NOT significant

How many sig figs are there in the numbers below?

0.0056 2

0.07089 4

0.000001

Rule #4: "Zeros on the right... sometimes significant" Zeros on the far right ARE significant if the number contains a decimal. Zeros on the far right are NOT significant if the number has no decimal point.

How many sig figs are there in the numbers below?

4.300×10^{-4} 4

200 1

0.010050 5

Rule #5: "Exact numbers" have an precise number of significant digits

Significant Digits when Adding or Subtracting

When you add or subtract numbers with different precision (decimal places), you must round your final answer to the least precise number place.

Example 1: A chemist adds 150.0 g sample to a beaker containing 0.507g. What is the total mass? 150.5

Significant Digits when Multiplying/Dividing

When you multiply or divide numbers with different precision, you must round your final answer to the least number of significant digits.

Example 2: What is the density of a 87.45 g metal sample with a volume of 2.0 cm³?

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

$$\rho = \frac{87.45 \text{ g}}{2.0 \text{ cm}^3} = 43.725 \text{ g/cm}^3$$

$$\rho = \frac{0.08745 \text{ kg}}{0.020 \text{ m}^3} = 4.4 \text{ kg/m}^3$$

Dimensional Analysis:

Molar mass 	$m = n \cdot M$ $m_{\text{mol}} \cdot \frac{\text{g/mol}}{1 \text{ mol}} = g$
Avogadro's number	$n \cdot 6.02 \times 10^{23} = \# \text{ of atoms}$
Concentration (molarity) 	$\frac{mol}{L} = \frac{mol/L}{1} = M$

Example 1: What is the mass in grams of an 8.4 mole sample of iron?

[470 g]

$$G: 8.4 \text{ mol} = n \quad \rightarrow m = 470 \text{ g}$$

$$R: M = ?$$

$$A: M = n \cdot M$$

$$S: M = 8.4 \text{ mol} \cdot 55.85 \text{ g/mol}$$

Example 2: A solution contains 0.42 moles of solute in 0.75 L. Calculate the molarity of the solution. [0.56M]

$$G: n = 0.42 \text{ mol}, V = 0.75 \text{ L}$$

$$R: C = ?$$

$$A: C = \frac{mol}{L}$$

$$S: C = \frac{0.42 \text{ mol}}{0.75 \text{ L}}$$

Example 3: A teaspoon of salt, NaCl has a mass of about 5.0 g. How many formula units are in a teaspoon of salt?

$$M = 22.99 \text{ g/mol} + 35.45 \text{ g/mol}$$

[5.2×10^{22} formula units]

$$G: M = 5.0 \text{ g}$$

$$\rightarrow M = 58.44 \text{ g/mol}$$

$$R: n = ?$$

$$A: n = \frac{m}{M}$$

$$S: n = \frac{5.0 \text{ g}}{58.44 \text{ g/mol}}$$

$$\rightarrow n = 0.08555 \text{ mol} \cdot 6.02 \times 10^{23}$$

$$n = 5.2 \times 10^{22} \text{ formula units}$$

CHECK YOUR UNDERSTANDING

- What is the mass of 5.00×10^{14} molecules of water? [$1.50 \times 10^{-8} \text{ g}$] $[5 \times 10^{14} \text{ molecules} \div 6.02 \times 10^{23}] \cdot 18.02 \text{ g/mol}$
- Convert 0.45 g of sodium hydroxide, NaOH to moles. [0.011 mol] $0.45 \text{ g} \div 40 \text{ g/mol} = 0.011 \text{ mol}$
- How many moles of hydrochloric acid, HCl, are present in 0.085 L of a 3.0 M solution? [0.26 mol]

$$\frac{0.45 \text{ g}}{40 \text{ g/mol}} = 0.011 \text{ mol}$$

$$(3) 0.085 \text{ L} \cdot 3.0 \text{ M}$$

$$= 0.26 \text{ mol}$$

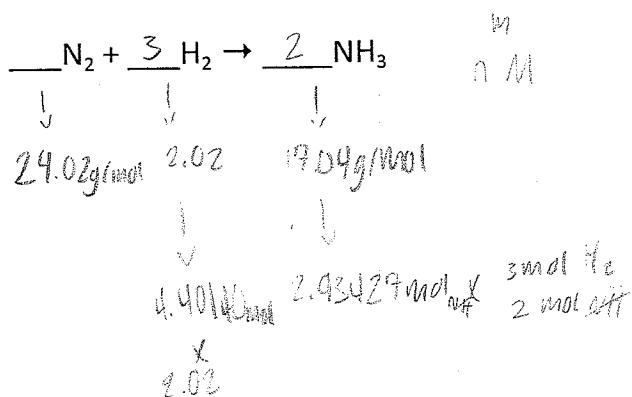
SCH4U Review: Stoichiometry

Example 1: What mass of hydrogen gas is needed to react completely with nitrogen gas to produce 50.0 g of ammonia gas? [8.90 g]

$$G: 50.0\text{g} = M_{\text{NH}_3}$$

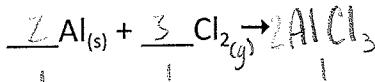
$$R: M_{\text{H}_2} = ?$$

$$S: M_{\text{H}_2} = 8.8\text{g}$$



Example 2: Aluminum reacts with chloride gas to form aluminum chloride. What mass of aluminum chloride can be produced from the reaction of 60.0g of aluminum metal? [297 g]

$$G: M_{\text{Al}} = 60.0\text{g}$$



$$\downarrow \quad \downarrow \quad \downarrow$$

$$26.98\text{g/mol} \quad 40.9\text{g/mol} \quad 132.33\text{g/mol}$$

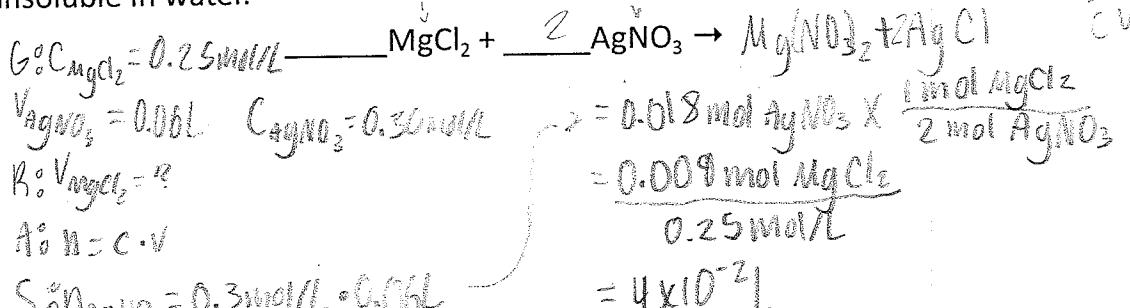
$$\downarrow \quad \downarrow$$

$$2.22386\text{mol} \quad 2.22386\text{mol}$$

$$\downarrow$$

$$246.5$$

Example 3: What is the minimum volume of 0.25 mol/L magnesium chloride, MgCl_2 , needed to precipitate all the silver ions in 60 mL of 0.30 mol/L silver nitrate, AgNO_3 ? Assume that silver chloride is completely insoluble in water.

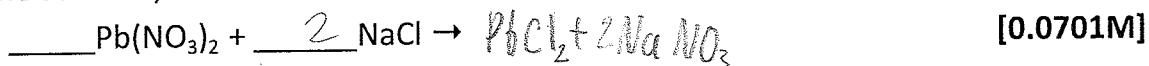


CHECK YOUR UNDERSTANDING

- How many molecules of sulfur trioxide will form when 2.50g of S_8 will react with excess oxygen:



- If 325 mL of lead (II) nitrate solution reacts with an excess solution of sodium chloride, 6.34 g of solid was filtered out of the solution, Find the molar concentration of the lead (II) nitrate solution.



*check notes for answers

n Limiting & Excess Reagents

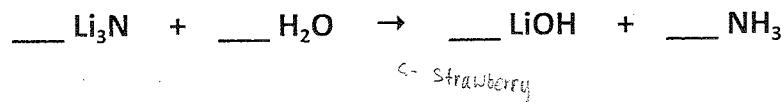
Example 1: What is the limiting reagent when 10.0g of $\text{Fe}_2(\text{SO}_4)_3$ reacts with an equal mass of NaOH? What mass of $\text{Fe}(\text{OH})_3$ is produced?

[$\text{Fe}(\text{SO}_4)_3$ is LR, 5.34g]



Example 2: In the following reaction, 4.87 g of lithium nitride reacts with 5.80 g of water. What is the limiting reactant? What mass of lithium hydroxide will be produced? What mass of excess reactant remains at the end of the reaction?

[H_2O , 7.71g, 1.13g]



CHECK YOUR UNDERSTANDING

1. 30.0 g of NH_4NO_3 and 50.0 g of Na_3PO_4 , determine the limiting reagent and the mass of ammonium phosphate that is produced? [NH_4NO_3 is LR, 18.6 g]
2. In a synthesis reaction, 62.0 g of P_4 and 4.00 g of H_2 react to produce PH_3 . What is the limiting reactant? [H_2] What mass of PH_3 will be produced? [44.9g]

TITRATION REVIEW

Term	Definition
Acid	
Base	
Titration	A laboratory technique used to determine the concentration of an unknown acid or base with a known base or acid.

Formulas: $pH = -\log[H^+]$ $pOH = -\log[OH^-]$ $pH + pOH = 14$ Logarithm Rule: #SD \rightarrow #DP

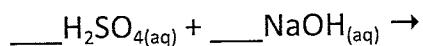
Example 1: What is the pH of a solution that contains 25 g of sulfuric acid dissolved in 1.34 L of water? [0.42]

Example 2: What is the pH of a solution containing 1.0×10^{-2} M NaOH? [12.00]

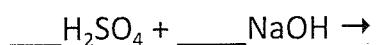
Example 3: 13.84 mL of sulfuric acid neutralizes 25.00 mL of a 0.100 mol/L solution of sodium hydroxide.

What is the concentration of sulfuric acid? Balance the equation first.

[0.0903M]



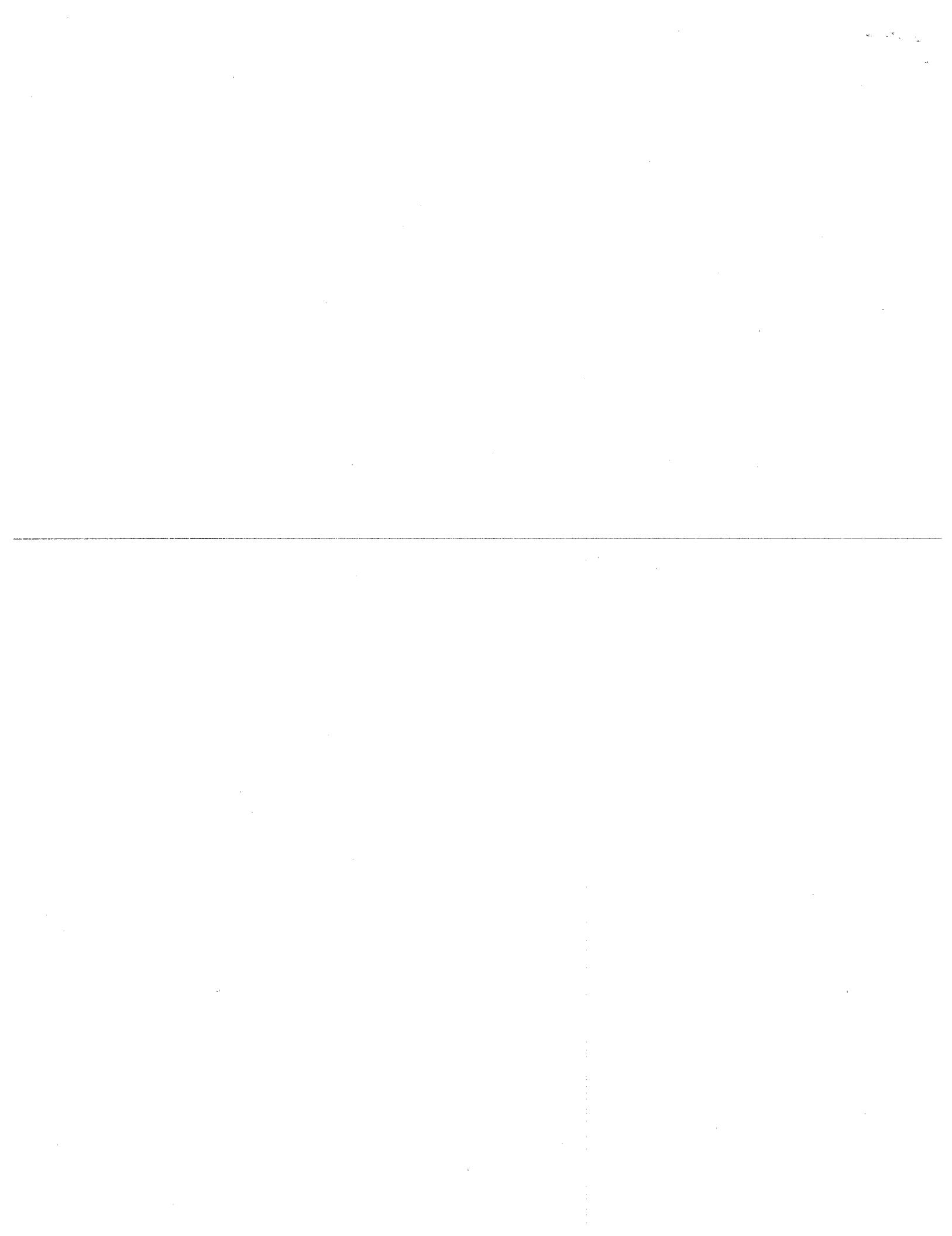
Example 2: What is the pH of a solution when you mix 95 mL of 0.100M H₂SO₄ to 45 mL of 0.200M NaOH?



[1.15]

CHECK YOUR UNDERSTANDING

1. What is the pH of a solution if you mix 75 mL of 0.125M HCl and 65 mL of 0.200M Mg(OH)₂? [pH = 13.07]
2. What is the pH of a solution if you mix 95 mL of 0.100M HCl and 45 mL of 0.200M NaOH? [pH = 2.45]
3. What is the pH of a solution if you mix 50.0 mL of 0.40M H₂SO₄ and 50.0 mL of 0.30M NaOH? [pH = 0.60]



GRADE 12 CHEMISTRY

Chapter 5: Thermochemistry

Topic	Homework	Learning Goals
5.1 Energy Changes in Chemical and Nuclear Rxns	p. 291 #2, 6, 9	<input type="checkbox"/> compare endothermic and exothermic reactions <input type="checkbox"/> write Thermochemical equations <input type="checkbox"/> draw an enthalpy diagram
5.2 Molar Enthalpy	p. 301 # 1- 3 p. 304 #1	<input type="checkbox"/> use stoichiometry ratios to determine ΔH
5.2 Calorimetry	p. 297 # 1, 2, 3 p. 301 #4 p. 305 #4 p. 306 # 1-7	<input type="checkbox"/> calculate the amount of heat transferred in a reaction using the formula $Q=mc\Delta T$ <input type="checkbox"/> calculate Molar enthalpy of a reaction using $\Delta H=Q/n$
5.3 Bond Energies	p. 312 #2, 3, 4 p. 313 #2, 3, 5, 8, 13	<input type="checkbox"/> use lewis structures of molecular compounds to calculate bond energy
5.4 Hess' Law	p. 317 # 1, 2, 3 p. 318 # 4, 5, 6, 7, 8	<input type="checkbox"/> apply Hess' Law to manipulate and combine different chemical equations to determine enthalpy change

Valentine's Day chemistry
COMPOUND INTEREST

Dopamine

Dopamine levels in the brain increase when you're in love, giving feelings of pleasure. People repeat behaviours that lead to dopamine release.

Serotonin

Studies have shown serotonin levels to be lower in people who are in love. They suggest these lower levels can lead to anxiety and obsession.

Adrenaline

Adrenaline, along with noradrenaline, is produced in stressful or exciting situations. It increases heart rate and contributes to the thrill of being in love.

Oxytocin

Oxytocin is sometimes called the 'cuddle hormone' and plays a key role in feelings of trust and attachment towards another person.

Endorphin

Endorphins are chemicals released by the body that act as natural painkillers. They are often released during physical exercise or in response to stress.

STRUCTURE KEY:

- Carbon
- O Oxygen
- N Nitrogen
- S Sulfur
- Hydrogen

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5.1 THERMOCHEMISTRY

Refer to pages 286 - 290 in your text and answer the following questions.

Thermochemistry:	
Energy:	
Work:	
Potential Energy:	
Kinetic Energy:	
Thermal Energy:	
Heat:	
Temperature:	
Law of Conservation of Energy:	
Chemical System:	
Open system:	
Closed system:	
Isolated system:	
Surroundings:	
Endothermic reaction:	
Exothermic reaction:	
Nuclear Reaction:	
Fusion:	
Fission:	

5.2 ENTHALPY CHANGE

Thermochemistry

- The study of energy changes involved in chemical and physical processes

Molar Enthalpy of a Reaction ΔH

- The amount of heat transferred during a reaction in kilojoules per mole of reactant.

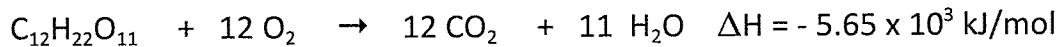
Exothermic Reaction	Endothermic Reaction
<ul style="list-style-type: none">A reaction that releases heat is exothermic and has a negative ΔHA candle burning release heatSurroundings get warmerHeat is a product $A + B \rightarrow C + D + \text{heat}$	<ul style="list-style-type: none">A reaction that absorbs heat is endothermic and has a positive ΔHMelting ice absorbs heatSurroundings get colderHeat is a reactant $A + B + \text{heat} \rightarrow C + D$

THREE WAYS OF REPRESENTING MOLAR ENTHALPY

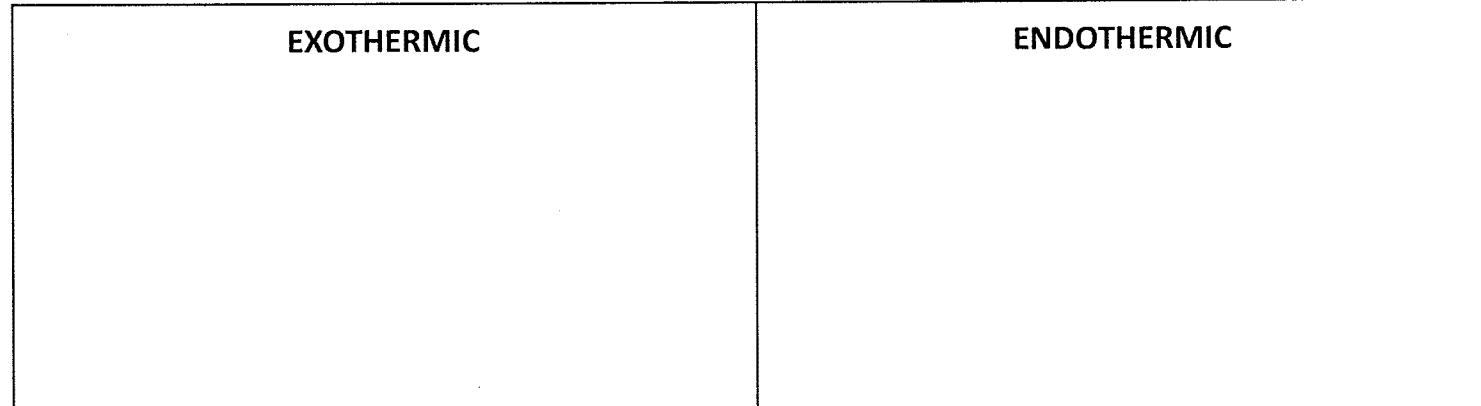
1. THERMOCHEMICAL EQUATIONS



2. INDICATE MOLAR ENTHALPY BESIDE THE CHEMICAL EQUATION



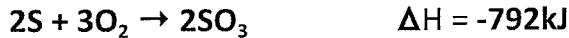
3. POTENTIAL ENERGY DIAGRAMS



NOTE: The arrow points in the direction that the energy is transferred.

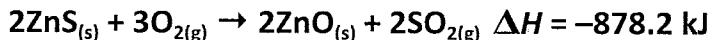
Example 1: Calculating Molar Enthalpy

How much heat is produced when 85.0 g of sulfur reacts according to the reaction below? [-1050 kJ]



THERMOCHEMICAL EQUATIONS & STOICHIOMETRY PRACTICE

1. Consider the following thermochemical equation:



a) How much heat is released when 3.000 mol $\text{ZnS}_{(\text{s})}$ reacts in excess oxygen? (-1317 kJ)

b) How much heat is released when 2.3×10^{-2} mol $\text{ZnS}_{(\text{s})}$ reacts in excess oxygen? (-1.0 x 10^1 kJ)

c) What is the enthalpy change when 223.9 g $\text{ZnS}_{(\text{s})}$ reacts in excess oxygen? (-1009 kJ)

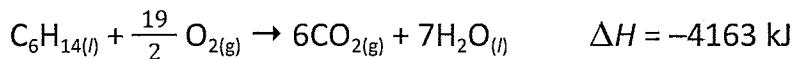
d) What is the enthalpy change when 0.96 g $\text{ZnO}_{(\text{s})}$ is produced? (-5.2 kJ)

2. Slaked lime ($\text{Ca(OH)}_{2(\text{s})}$) is produced when lime (calcium oxide, $\text{CaO}_{(\text{s})}$) reacts with liquid water. 65.2 kJ of heat is released for each mol of Ca(OH)_2 that is produced. (Note: 1 kg = 1000g)

a) Write a thermochemical equation for the reaction.

b) What is the enthalpy change when 523 kg of lime reacts with excess water? (-6.08 x 10^5 kJ)

3. The following reaction represents the complete combustion of hexane, $\text{C}_6\text{H}_{14(\text{l})}$, at SATP.



a) If 0.537 mol of carbon dioxide is produced in the reaction represented by the equation above, how much heat is released by the reaction? (-373 kJ)

b) If 25.0 kg of hexane is burned in sufficient oxygen, how much heat will be released? (-1.21x 10^6 kJ)

c) What mass of hexane is required to produce 1.0×10^5 kJ of heat by complete combustion?

(2.1x 10^3 g)

5.2 CALORIMETRY AND ENTHALPY

Calorimetry

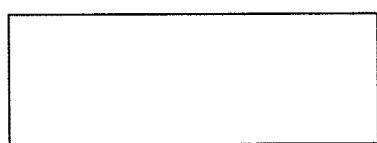
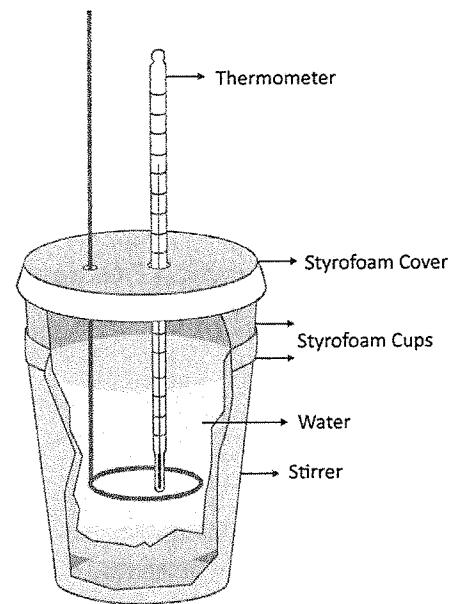
- the process of measuring energy changes in a chemical system

Calorimeter

- is used to measure the heat absorbed or released in a reaction by measuring the temperature change of an insulated mass of water
- assume that heat is transferred to/from the water (aqueous surroundings) only

Specific Heat Capacity, c (units = J/g°C)

- The amount of energy needed to raise the temperature of one gram of substance 1°C
- the specific heat capacity can be used to determine the amount of energy required or released during a physical or chemical change



where

Q = amount of heat (Joules)

m = mass (g)

c = specific heat capacity (J/g·°C)

ΔT = change in temperature (°C) $T_2 - T_1$

Example 1: Solar energy can preheat cold water for domestic hot-water tanks. What quantity of heat is obtained from solar energy if 100.0 kg of water is pre-heated from 10.0°C to 45.0°C?

USING A CALORIMETER TO OBTAIN ΔH VALUES

- no heat is transferred between the calorimeter and the outside environment (all heat is absorbed by the water hence the mass of the water is used in the equation)
- any heat absorbed or released by a Styrofoam cup calorimeter is negligible ($c_{cup} = 0$ J/g·°C)
- a dilute aqueous solution is assumed to have a density and specific heat capacity equal to that of pure water ($d = 1.00$ g/ml, $c = 4.184$ J/g·°C or 4.184 kJ/kg·°C)

CASE 1: SOLID → H₂O

In a laboratory investigation into the reaction:



A researcher adds a 261 g sample of barium nitrate to 2.0 L of potassium sulfate solution in a polystyrene calorimeter. As the barium nitrate dissolves, a precipitate is immediately formed. T₁ = 26.0 °C, T₂ = 29.1 °C Calculate the molar enthalpy of the reaction of barium nitrate.

STEP 1: Use q = mcΔT to determine heat change

STEP 2: Change the sign of q

$$q_{\text{system}} = -q_{\text{surroundings}}$$

$$\Delta H_{\text{system}} = -q_{\text{surroundings}}$$

The amount of heat released or absorbed by the reaction will be the opposite to the surroundings. We need to convert to surroundings because enthalpy depends on the heat of surroundings.

STEP 3: Calculate moles. Since we are calculating in terms of barium nitrate we will use the molar mass to find moles.

STEP 4: Find Molar Enthalpy $\Delta H = \frac{-q}{n}$

CASE 2: TWO LIQUIDS

Given the following reaction:



50.0 mL of 0.300 mol/L CuSO₄ solution is mixed with an equal volume of NaOH. The initial temperature of both solutions is 21.4°C. After mixing the solutions in the coffee-cup calorimeter, the highest temperature that is reached is 24.6°C.

- a) Determine the enthalpy change, ΔH of the reaction.
- b) Then write a thermochemical equation.

Bomb Calorimeter:

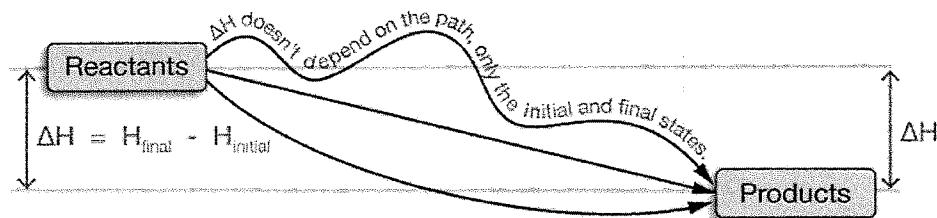
- Most experiments can be done at constant pressure in a coffee cup calorimeter or a laboratory calorimeter (see Figure 2 on page 293).
- The molar enthalpy of fuels however is calculated by burning the fuel in a bomb calorimeter.
- This type of calorimeter is a very rigid, tightly sealed container which prevents the gases from escaping.
- The heat released by the combustion will be absorbed by the water and the calorimeter.

CASE 3: BOMB CALORIMETER

A bomb calorimeter composed of steel with a mass of 322g is filled with 225g of water. The original temperature of the water and calorimeter is 10.6°C. When 1.02 g of ethanol is burned, the temperature rises to 38.4°C. Calculate the molar enthalpy of combustion of ethanol. Specific heat of steel = 0.44 J/g·°C.

5.4 HESS' LAW

The enthalpy change of a physical or chemical process depends only on the initial and final conditions of the process; The enthalpy change of the overall process is the sum of the enthalpy changes of its individual step. Thus multi-step equations can be added to determine the overall enthalpy.



RULES:

1. If a reaction must be reversed, the sign of the ΔH for that reaction must be changed.
2. When adjusting molar ratios by multiplying, all coefficients in that equation must be multiplied by the same factor, as must the ΔH value for that reaction.

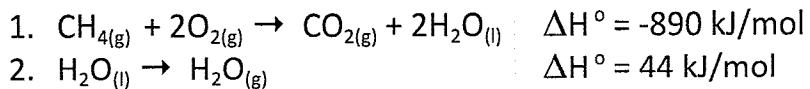
Example 1: Carbon solid reacts with water vapour to form carbon monoxide gas and hydrogen gas in the following reaction.



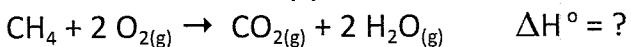
Use the following reference equations to calculate the ΔH value for this reaction.

1. $C_{(s)} + O_2_{(g)} \rightarrow CO_2_{(g)}$ $\Delta H = -393 \text{ kJ/mol of C}$
2. $CO_{(g)} + \frac{1}{2} O_2_{(g)} \rightarrow CO_2_{(g)}$ $\Delta H = -283 \text{ kJ/mol of CO}$
3. $\frac{1}{2} O_2_{(g)} + H_2_{(g)} \rightarrow H_2O_{(g)}$ $\Delta H = -242 \text{ kJ/mol of H}_2$

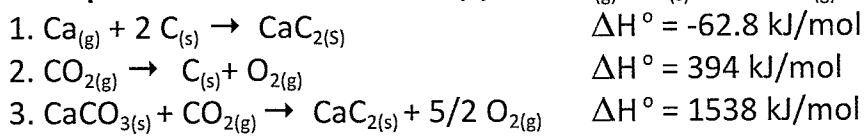
Example 2: From the following data,



Calculate the enthalpy of the reaction:

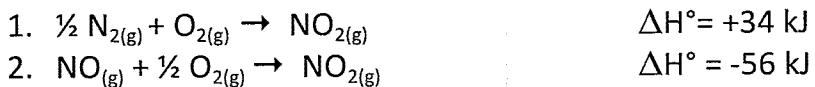


Example 3: Calculate the enthalpy for: $\text{Ca}_{(g)} + \text{C}_{(s)} + 3/2 \text{O}_{2(g)} \rightarrow \text{CaCO}_{3(s)}$ using the information below



Example 4:

Given:



Calculate the standard enthalpy change for the formation of two moles of nitrogen monoxide from its elements.

5.3 BOND ENERGIES

Bond breaking is an endothermic process and bond forming is an exothermic process, the enthalpy of a reaction can be calculated using average bond energies.

$$\Delta H = \sum E \text{ bonds broken} - \sum E \text{ bonds formed}$$

Bond energy can be used to calculate ΔH but this is only an estimate because bond energies are average values from the same bond in a number of different molecules. See Table 1 on page 307.

Before summing, the bond energies must be multiplied by the number of that bond present AND the coefficient from the balanced chemical equation.

When writing the balanced equation, the structures of each species must be drawn to determine the types of bonds present.

Multiple bonds (that is, double or triple bonds) have larger bond energies than single bonds. Also, the length of the bond shortens as the number of bonds increases.

Example 1: Use bond energies to determine heat of reaction for the complete combustion of methane.

Example 2: Hydrogen gas reacts with fluorine gas to form hydrogen fluoride. Use bond energies to calculate the heat of this reaction per one mole of hydrogen fluoride.



Bond Energy

Chem Worksheet 16-2

Name _____

Period _____

Bond energy is defined as the amount of energy required to break a bond. These values are positive, indicating that bond breaking is endothermic. Bond energies are reported in kilojoules per mole (kJ/mol). The energy for *breaking* a hydrogen-hydrogen bond is 436 kJ/mol so when a hydrogen-hydrogen bond is *formed* the process releases 436 kJ/mol.

In a chemical reaction several bonds are broken and formed. For example in the reaction below a hydrogen-hydrogen bond is broken and a fluorine-fluorine bond is broken. Two hydrogen-fluorine bonds are formed. The overall energy change for this process is calculated below.



Bond energy: 436 kJ/mol

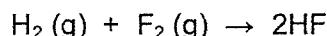


Bond Energy: 155 kJ/mol



Bond Energy: 567 kJ/mol

example



$$\Delta H_{rxn} = [\text{energy used for breaking bonds}] - [\text{energy formed in making bonds}]$$

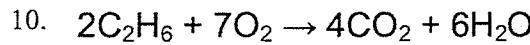
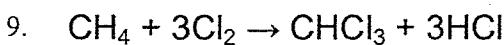
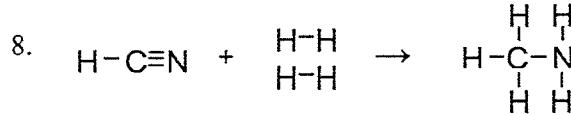
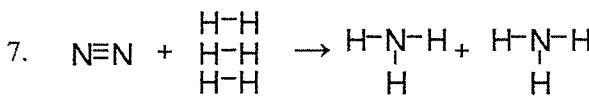
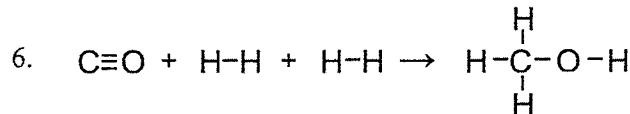
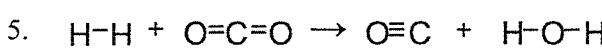
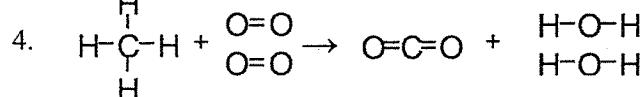
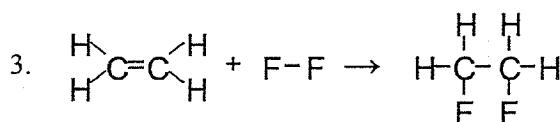
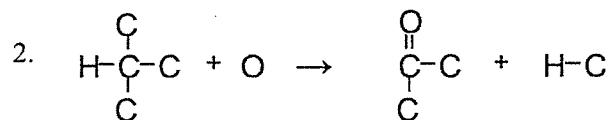
$$\Delta H_{rxn} = [436 \text{ kJ/mol} + 155 \text{ kJ/mol}] - [2(567 \text{ kJ/mol})]$$

$$\Delta H_{rxn} = -543 \text{ kJ/mol}$$

Average Bond Energies (kJ/mol)

H-H	436 kJ/mol	C-H	413 kJ/mol	C=C	614 kJ/mol
H-Cl	431 kJ/mol	C-C	348 kJ/mol	C≡C	839 kJ/mol
H-F	567 kJ/mol	C-N	293 kJ/mol	C=O	799 kJ/mol
N-H	391 kJ/mol	C-O	358 kJ/mol	O=O	495 kJ/mol
N-O	201 kJ/mol	C-F	485 kJ/mol	C≡O	1072 kJ/mol
O-H	463 kJ/mol	C-Cl	328 kJ/mol	N=N	418 kJ/mol
O-O	146 kJ/mol	C-S	259 kJ/mol	N≡N	941 kJ/mol
F-F	155 kJ/mol	Cl-Cl	242 kJ/mol	C≡N	891 kJ/mol

Estimate the enthalpy change (ΔH_{rxn}) of the following reactions using the bond energies above.



5.3 Average Bond Energies (kJ/mol)

Single Bonds			Multiple Bonds
H-H 432	N-H 391	I-I 149	C=C 614
H-F 565	N-N 160	I-Cl 208	C≡C 839
H-Cl 427	N-F 272	I-Br 175	O=O 495
H-Br 363	N-Cl 200	S-H 347	C=O* 745
H-I 295	N-Br 243	S-F 327	C≡O 1072
C-H 413	N-O 201	S-Cl 253	N=O 607
C-C 347	O-H 467	S-Br 218	N=N 418
C-N 305	O-O 146	S-S 266	N≡N 941
C-O 358	O-F 190	Si-Si 340	C≡N 891
C-F 485	O-Cl 203	Si-H 393	C=N 615
C-Cl 339	O-I 234	Si-C 360	
C-Br 276	F-F 154	Si-O 452	
C-I 240	F-Cl 253		
C-S 259	F-Br 237		
	Cl-Cl 239		
	Cl-Br 218		
	Br-Br 193		

*C=O in CO₂(g) = 799

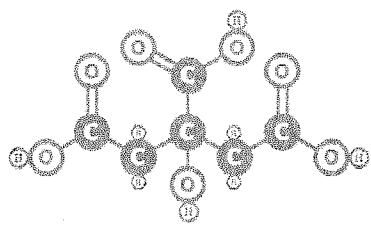
Chapter 8A Acids & Bases

Homework Schedule

Section	Homework	Learning Goals
8.1 The Nature of Acids and Bases	p. 492 #1-2, p. 493 #1, p. 494 #1-9	<input type="checkbox"/> Identify Bronsted-Lowry acid/base pairs and conjugate acids/bases in chemical reactions. <input type="checkbox"/> Distinguish between strong and weak acids/bases, and understand their ionization and equilibrium expressions
8.2 pH & pOH Calculations	p. 502 #1-2, p. 505 #1-3, p. 508 #1-4, p. 509 #1-9	<input type="checkbox"/> Understand ionization and equilibrium constants for acids and bases <input type="checkbox"/> Calculate pH and pOH given concentration of various solutions.
8.4 & 8.5 Calculations of Weak Acids and Bases	p. 513 #1-2, p. 516 #1-2, p. 520 #1-2, p. 521 #1-2, p. 524 #1-2, p. 527 #1-2, p. 529 #1-3	<input type="checkbox"/> Distinguish between when to use stoichiometry or ICE charts for strong or weak acids and bases <input type="checkbox"/> Calculate the K_a/K_b and % ionization of weak acids and bases <input type="checkbox"/> Use K_a/K_b to solve for the pH of weak acids or bases

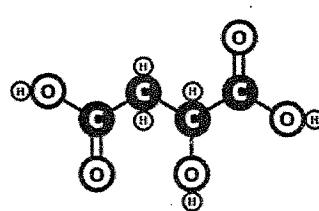
A GUIDE TO COMMON FRUIT ACIDS

Most people probably know that lemons and other citrus fruits contain citric acid – but it's just one of a number of different organic acids that can be found in fruits. Here we look at a number of the most common acids, and the various fruits that they are found in.



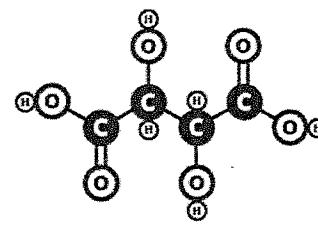
CITRIC ACID

The main acid in citrus fruits is, unsurprisingly, citric acid. Lemons and limes have particularly high levels of this compound. It is also the main acid in a number of berry fruits, including strawberries, raspberries and gooseberries.



MALIC ACID

Malic acid is the main acid in most stone fruits such as cherries, apricots, peaches, and nectarines. It's also found in high amounts in apples, and in lower amounts in bananas. Though watermelons have a low acid content, their principal acid is also malic acid.

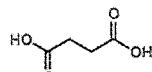


TARTARIC ACID

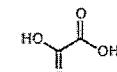
Tartaric acid is the principle acid in fewer fruits than citric and malic acid. However, it is the main acid in grapes, which also contain malic acid. Red grapes have higher levels of tartaric acid. The main acid of avocado and tamarind is also tartaric acid.

OTHER ORGANIC ACIDS

Citric, malic, and tartaric acids are not the only organic acids present in fruit – a number of other acids are also present, albeit in significantly smaller quantities. To the right, a small selection of these compounds are shown, along with a brief note of some of the fruits in which they're often found.



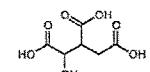
SUCCINIC ACID



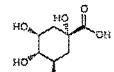
OXALIC ACID



BENZOIC ACID



LACTIC ACID



QUINIC ACID

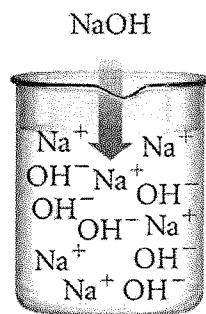
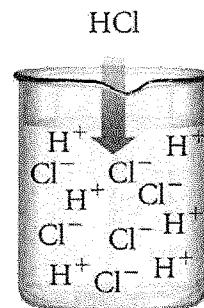


8.1 THE NATURE OF ACIDS AND BASES

THE ARRHENIUS THEORY

Arrhenius Acid

- a substance that dissociates in water to produce _____ ions
e.g.



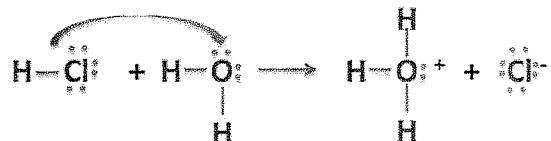
Arrhenius Base

- A substance that dissociates in water to produce _____ ions
e.g.

Limitations of the Arrhenius Theory:

- Rather than just dissociating in water ...

H^+ attaches itself to a water molecule in an aqueous solution, forming the _____ ion,



- Cannot explain the basic properties of ammonia, NH_3 as it does not have an OH^- ion

THE BRONSTED-LOWRY THEORY

Bronsted-Lowry Acid

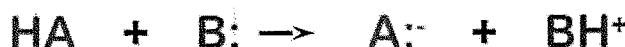
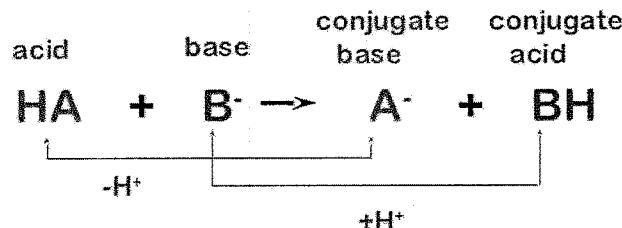
- a substance from which a proton (H^+) can be removed (_____ H^+)

Bronsted-Lowry Base

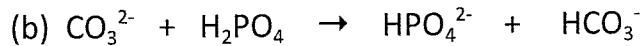
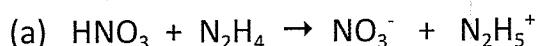
- a substance that can remove a proton (H^+) from an acid (_____ H^+)

The Bronsted-Lowry theory doesn't go against the Arrhenius theory – it just adds to it

Any negative ion (anion) can be a Bronsted-Lowry base, not just OH^-



e.g. Identify the Bronsted-Lowry acid/base pairs in the following reactions:



e.g. What is the conjugate **base** of HSO_4^- ?

What is the conjugate **acid** of HSO_4^- ?

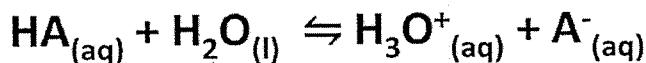
** notice that HSO_4^- can both accept and donate a proton

Amphoteric Substance

- a substance which can act as either an acid or a base
- will contain both a H^+ and an overall negative charge or lone pair

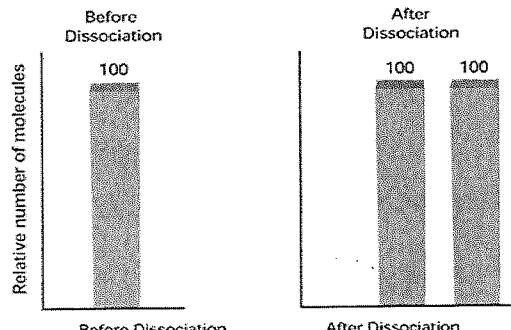
Strong vs Weak Acids

Acids will react with water ...



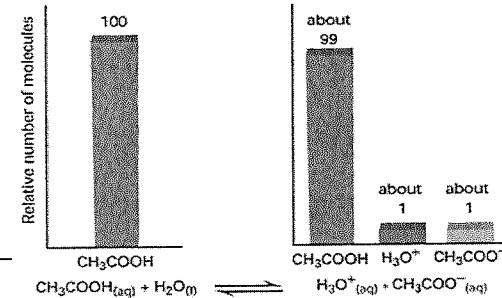
Strong acids

- have 100% ionization,
- =
- (_____ arrow, use _____)



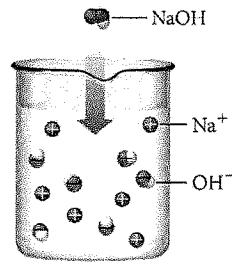
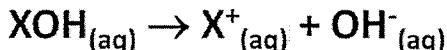
Weak acids

- only partially ionize
- small K_a , equilibrium lies far to the left
- <<
- (_____ arrow, use _____)



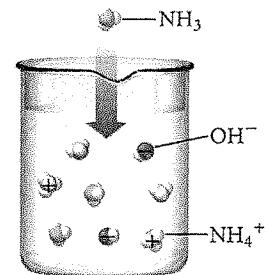
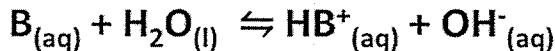
Strong bases

- have 100% ionization, producing OH⁻
- no K_b value
- =
- (_____ arrow, use _____)



Weak bases

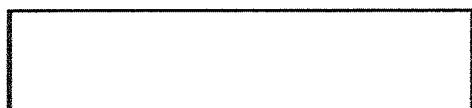
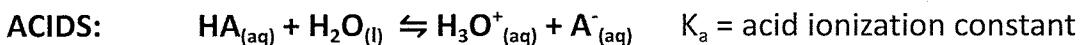
- react with water to produce OH⁻
- small K_b value
- <<
- (_____ arrow, use _____)



<u>Strong Acids</u>	<u>Weak Acids</u>
HCl	all others (H___)
HBr	
HI	
H ₂ SO ₄	
HNO ₃	
HClO ₄	

<u>Strong Bases</u>	<u>Weak Bases</u>
metal + OH ⁻	do not contain OH ⁻
NaOH	NH ₃
KOH	CH ₃ NH ₂
Ca(OH) ₂	C ₅ H ₅ N
etc	etc
(Arrhenius)	(Bronsted-Lowry)

Acid-Base Equilibrium



Note: H₂O_(l) is not included in expression

The smaller the K_a, the _____ the acid and the _____ the acid ionizes



Note: H₂O_(l) is not included in expression

The smaller the K_b, the _____ the base and the _____ the base ionizes

Ex. Write the K_a expression for acetic acid,
CH₃COOH

Ex. Write the K_b expression for ammonia,
NH₃

8.2 WATER AS AN ACID OR A BASE

Pure water ionizes according to the following reaction: _____

Equilibrium constant for the dissociation of water: _____

THE RELATIONSHIP BETWEEN K_a and K_b

For a conjugate acid-base pair

Ex) What is the K_b value for F^- if the K_a value of HF is 6.6×10^{-4} ?

The **larger** the K_a , the **smaller** the K_b . That is, the **stronger** the acid, the **weaker** the conjugate base.

A _____ - acid or base will produce a _____ with _____ acid/base properties

$$\text{HCl}_{(\text{aq})} + \text{H}_2\text{O}_{(\ell)} \rightleftharpoons \text{H}_3\text{O}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$$

A _____ acid or base will produce _____ that has _____ acid/base properties

$$\text{HClO}_{(\text{aq})} + \text{H}_2\text{O}_{(\ell)} \rightleftharpoons \text{H}_3\text{O}^+_{(\text{aq})} + \text{ClO}^-_{(\text{aq})}$$

POLYPROTIC ACIDS

- acids that have more than one proton
- only one proton is transferred at a time

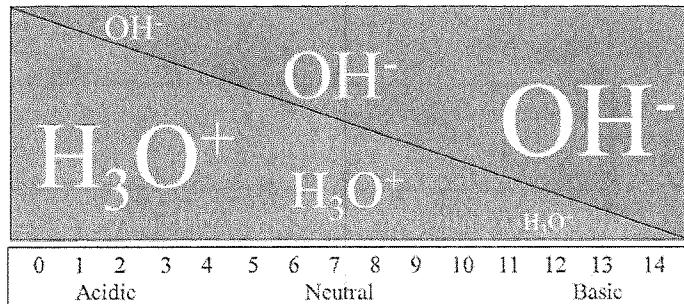
Ex) $\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons$

Acid-Base Relationships

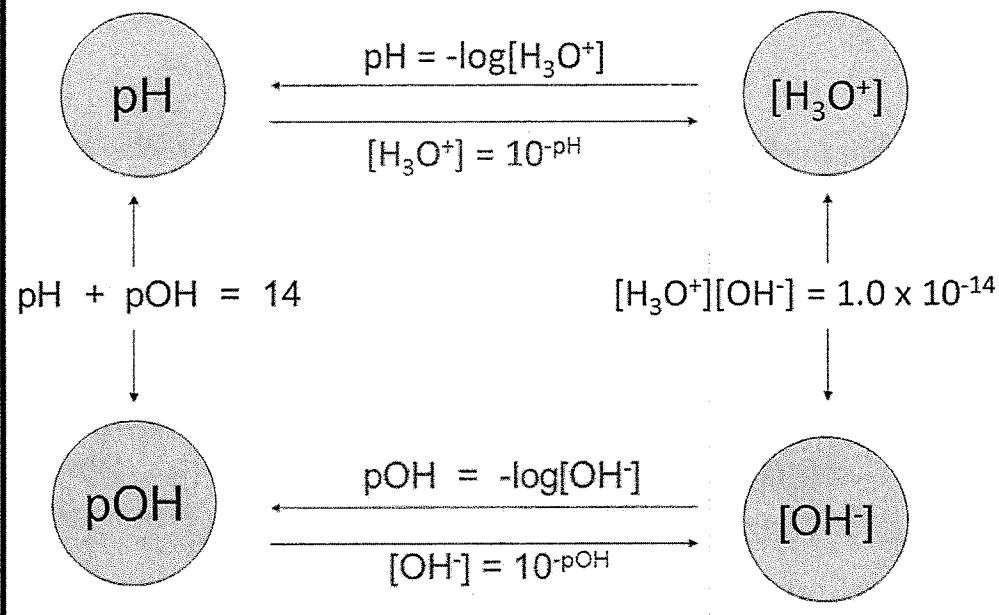
$$[\text{H}_3\text{O}^+] > [\text{OH}^-] \quad [\text{H}_3\text{O}^+] = [\text{OH}^-] \quad [\text{H}_3\text{O}^+] < [\text{OH}^-]$$

pH and pOH

As $[\text{H}_3\text{O}^+]$ rises, $[\text{OH}^-]$ falls...



pH Calculations



e.g. A liquid shampoo has a $[\text{OH}^-]$ of 6.8×10^{-5} mol/L

(a) What is $[\text{H}_3\text{O}^?]$

(b) What is the pH and pOH of the shampoo?

8.4 CALCULATIONS INVOLVING ACIDIC SOLUTIONS

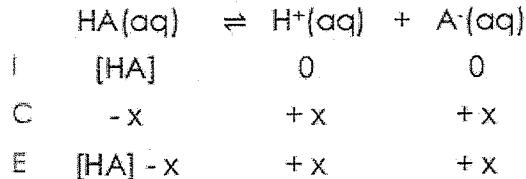
Strong Acid

Strong acids dissociate completely into ions in aqueous solution.

$$[\text{H}^+] = [\text{HA}]$$

Weak Acid

$$K_a = \frac{x^2}{([\text{HA}] - x)}$$



Calculations Involving STRONG Acids: (STOICHIOMETRY!!)

e.g. Find the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in 0.50 mol/L nitric acid (a strong acid). What is the pH?

Calculations Involving WEAK Acids: (ICE CHARTS!!)

e.g. A 0.200 mol/L solution of a weak acid HA, is 5% ionization. What is the ionization constant, K_a ?

e.g. Propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}$ is a weak acid. A 0.10 mol/L solution has a pH of 2.96. What is the K_a for this acid? What is the % ionization?

e.g. What is the pH of a 0.025 mol/L solution of formic acid, HCOOH ? $K_a = 1.8 \times 10^{-4}$

8.5 CALCULATIONS INVOLVING BASIC SOLUTIONS

Strong Base	Weak Base	$K_b = \frac{x^2}{([B] - x)}$
Strong bases dissociate completely into ions in aqueous solution.	$B + H_2O \rightleftharpoons BH^+ + OH^-$	
$[OH^-] = [B]$	I [B] - 0 0	
	C -x - +x +x	
	E [B] - x - x x	

Calculations Involving STRONG Bases:

(STOICHIOMETRY!!)

Ex) Find the $[OH^-]$, $[H_3O^+]$, pOH and pH of a 0.16 M solution of barium hydroxide.

Calculations Involving WEAK Bases:

(ICE CHARTS!!)

Ex) Pyridine, C_5H_5N is a weak base. A 0.125 M solution has a pH of 9.10. Calculate the K_b .

Ex) Ammonia has a K_b value of 1.8×10^{-5} . What is the pH of a 0.105 M solution of ammonia?

B5 K_a and K_b for Common Acids and Weak Bases

Table 1 Monoprotic Acids

Name	Formula of Acid	Formula of Conjugate Base	Equilibrium Constant, K_a
perchloric acid	$\text{HClO}_4\text{(aq)}$	$\text{ClO}_4^-\text{(aq)}$	very large
hydroiodic acid	$\text{HI}\text{(aq)}$	$\text{I}^-\text{(aq)}$	very large
hydrobromic acid	$\text{HBr}\text{(aq)}$	$\text{Br}^-\text{(aq)}$	very large
hydrochloric acid	$\text{HCl}\text{(aq)}$	$\text{Cl}^-\text{(aq)}$	very large
nitric acid	$\text{HNO}_3\text{(aq)}$	$\text{NO}_3^-\text{(aq)}$	very large
hydronium ion	$\text{H}_3\text{O}^+\text{(aq)}$	$\text{H}_2\text{O(l)}$	1.0
hydrogen sulfate ion	$\text{HSO}_4^-\text{(aq)}$	$\text{SO}_4^{2-}\text{(aq)}$	1.2×10^{-2}
chlorous acid	$\text{HClO}_2\text{(aq)}$	$\text{ClO}_2^-\text{(aq)}$	1.2×10^{-2}
iron(III) ion	$\text{Fe}(\text{H}_2\text{O})_6^{3+}\text{(aq)}$	$\text{Fe}(\text{H}_2\text{O})_5(\text{OH})^{2+}\text{(aq)}$	1.5×10^{-3}
citric acid	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7\text{(aq)}$	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-\text{(aq)}$	7.4×10^{-5}
hydrofluoric acid	$\text{HF}\text{(aq)}$	$\text{F}^-\text{(aq)}$	6.6×10^{-4}
nitrous acid	$\text{HNO}_2\text{(aq)}$	$\text{NO}_2^-\text{(aq)}$	4.6×10^{-4}
hydrogen cyanate	HOCN	$\text{OCN}^-\text{(aq)}$	3.5×10^{-4}
methanoic acid	HCHO_2 $\text{HCOOH}\text{(aq)}$	$\text{CHO}_2^-\text{(aq)}$	1.8×10^{-4}
chromium(III) ion	$\text{Cr}(\text{H}_2\text{O})_6^{3+}\text{(aq)}$	$\text{Cr}(\text{H}_2\text{O})_5(\text{OH})^{2+}\text{(aq)}$	1.0×10^{-4}
methyl orange	$\text{HMO}\text{(aq)}$	$\text{MO}^-\text{(aq)}$	$\sim 10^{-4}$
benzoic acid	$\text{HC}_6\text{H}_5\text{O}_2\text{(aq)}$ $\text{C}_6\text{H}_5\text{COOH}\text{(aq)}$	$\text{C}_6\text{H}_5\text{O}_2^-\text{(aq)}$ $\text{C}_6\text{H}_5\text{COO}^-\text{(aq)}$	6.3×10^{-5}
ethanoic (acetic) acid	$\text{HC}_2\text{H}_3\text{O}_2\text{(aq)}$ $\text{CH}_3\text{COOH}\text{(aq)}$	$\text{C}_2\text{H}_3\text{O}_2^-\text{(aq)}$	1.8×10^{-5}
aluminum ion	$\text{Al}(\text{H}_2\text{O})_6^{3+}\text{(aq)}$	$\text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+}\text{(aq)}$	9.8×10^{-6}
bromothymol blue	$\text{HBB}\text{(aq)}$	$\text{Bb}^-\text{(aq)}$	$\sim 10^{-7}$
hypochlorous acid	$\text{HClO}\text{(aq)}$	$\text{ClO}^-\text{(aq)}$	3.5×10^{-8}
phenolphthalein	$\text{HPh}\text{(aq)}$	$\text{Ph}^-\text{(aq)}$	$\sim 10^{-10}$
hydrocyanic acid	$\text{HCN}\text{(aq)}$	$\text{CN}^-\text{(aq)}$	6.2×10^{-10}

Chapter 8 Titrations

Homework Schedule

Section	Homework	Learning Goals
8.6 Salts	p.534 #1-2, p. 536 #1, p. 538 #1-2, p. 539 #1-5	<input type="checkbox"/> I can determine the pH of a solution containing a salt
8.7 Titrations	p. 547 #1-2, p. 554 #1-2, p. 557 #1-8	<input type="checkbox"/> I can solve for pH for titrations with strong-strong <input type="checkbox"/> I can solve for pH for titrations with strong-weak
8.8 Buffers	p. 565 # 1 – 3, p. 567 # 1 – 8 Practice Problems	<input type="checkbox"/> I can identify a buffer solution <input type="checkbox"/> I can solve for pH of buffers

CHEMISTRY TECHNIQUES: TITRATION

Used to determine the concentration of a particular solution, by measuring how much of a solution of known concentration reacts with a known volume of it.

SALTING-OUT

The pipette comes in a range of fixed volumes, or in graduated form, depending on the volume you need to measure. Burettes usually hold 50 cm³ of solution.

ASSESSING REACTANTS

Pipettes should never be filled direct from stock bottles of solution as this could contaminate them. Instead, fill them from some of the solution poured into a beaker. They calibrate to allow for a small amount of solution remaining in the tip after the correct volume has been delivered, so this remainder shouldn't be forced out.

FILLING THE BURETTE

As with the pipette, the burette should be rinsed with distilled water then the solution it is to be filled with before filling. It is necessary that the area under the tap also contains solution as the burette's scale includes this volume. Without doing this, the volume recorded for titrations would be higher than the actual value.

USING INDICATORS

PHENOLPHTHALEIN

METHYL ORANGE

In acid-base titrations, a range of indicators can be used. These are solutions which change colour at a specific pH, and can be used to precisely identify when the neutralisation reaction is complete (the end point). Different indicators are suitable for different acid-base combinations.

CARRYING OUT A TITRATION

TEST TITRANT →

20.00 cm³ • 20.15 cm³ • 20.00 cm³ • 20.10 cm³

* concordant titre values, used to calculate average

Average titre = (20.15 cm³ + 20.00 cm³) / 2 = 20.075 cm³

To carry out the titration, the top of the burette is opened to allow the solution inside to flow into a known volume of the solution in the conical flask. The amount of solution from the burette required to reach the end point is recorded. A rough titration is usually followed by more accurate runs. Multiple titrations are carried out until concordant titres are obtained (within 0.10 cm³ of each other).

CARRYING OUT CALCULATIONS

$n = CV$

Assuming that the concentration of the solution in the conical flask is 1 M.

Calculate number of moles of solution added from the burette.

Determine the number of moles in the conical flask using the equation for the reaction and reacting ratios.

Calculate the concentration of the solution in the conical flask by rearranging the equation ($c = n / V$).

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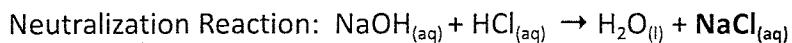
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8.6 ACID-BASE PROPERTIES OF SALT SOLUTIONS

Salts

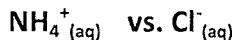
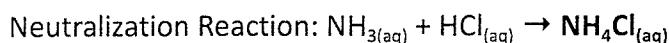
- an ionic compound containing a cation (+ ion) and an anion (- ion)
- the cation comes from the _____ and the anion comes from the _____
- can be _____, _____ or _____
- the pH is dependent upon whether the ions are conjugates of strong or weak acids and bases

Salts that Produce Neutral Solutions

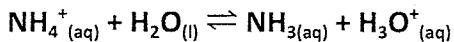


- Cation (i.e., Na^+) is the conjugate of a _____ base (i.e., NaOH), and therefore has _____ acidic properties.
- Anion is the conjugate of a _____ acid (i.e., HCl) and therefore has _____ basic properties.
- Neither of the ions act as an acid or a base and thus the pH is _____
- a salt produced in the reaction of a _____ and _____ will be _____

Salts that Produce Acidic Solutions

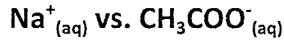
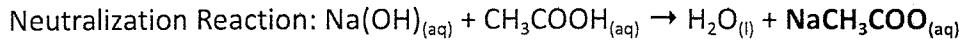


- Cation (NH_4^+) is the _____ conjugate acid of a _____ base (NH_3). This conjugate acid would thus react with water to produce H_3O^+ .

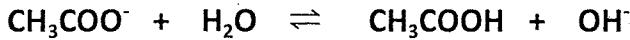


- Anion (Cl^-) is the conjugate of a _____ acid (i.e., HCl), and therefore would have negligible basic properties.
- a salt produced in the reaction of a _____ and _____ will be _____

Salts that Produce Basic Solutions



- Anion (CH_3COO^-) is the _____ conjugate base of a _____ acid (CH_3COOH).
- This conjugate base would thus react with water to produce OH^- .



- Cation is the conjugate of a _____ base (i.e., NaOH), and therefore will have negligible acidic properties.
- a salt produced in the reaction of a _____ and _____ will be _____

What if both cation and anion come from a weak acid and a weak base like NH_4HCO_3 ?

Steps for predicting the pH of a salt

1. For a 0.10 mol/L solution of NaNO_2 . Predict if it will be acidic, basic or neutral. Calculate the pH

2. Is the solution of KH_2PO_4 acidic or basic?

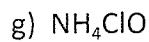
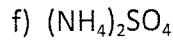
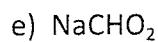
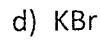
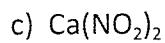
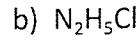
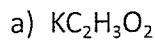
Table 2 Polyprotic Acids

Name	Formula of Acid	Formula of Conjugate Base	Equilibrium Constant		
			K_{a_1}	K_{a_2}	K_{a_3}
phosphoric acid	$\text{H}_3\text{PO}_4(\text{aq})$	$\text{H}_2\text{PO}_4^-(\text{aq})$	7.1×10^{-3}	6.3×10^{-8}	4.2×10^{-13}

3. Calculate the mass of NH_4Br that has to be dissolved in 1.00 L of water to have a solution with a pH of 5.15. K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$

Try it Yourself! Salts Practice Problems

1. Predict whether the solutions are acidic, basic or neutral. If not neutral, write the equation for the reaction that causes the solution to be acidic or basic.



8.7 ACID-BASE TITRATIONS

Titration

- used to determine the unknown concentration of an acid or a base

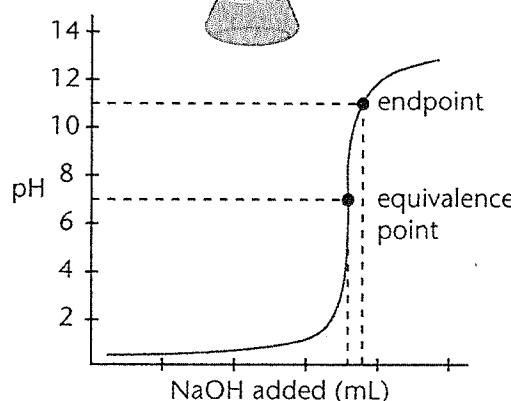


Equivalence point

- the point where the acid and base are _____ in _____
- on a titration curve, it is the midpoint on the steep rise

Endpoint

- occurs when the _____ changes _____
- as the pH changes rapidly near the equivalence point, the difference between end point and equivalence point is negligible



Titrations Involving Strong Acids and Strong Bases

e.g. a) What volume of 0.100 M NaOH is required to reach the equivalence point when titrating 50.0 mL of 0.200 M HNO₃?

b) What is the pH at the equivalence point?

- At the equivalence point all the acid and base have been _____.
- Since this is a titration between a _____ acid and a _____ base, the salt produced (NaNO₃) will have no effect on pH.
- The pH at the equivalence point will be _____.

c) Calculate the pH of HNO₃ before any base is added.

d) Calculate the pH of the solution after 50.0 mL of 0.100 M NaOH is added to 50.0 mL of 0.200 M HNO₃.

Titration Problems Involving Weak Acids or Bases are done using an ICE Table.

In a titration, 20.00 ml of 0.300 mol/L $\text{CH}_3\text{COOH}_{(\text{aq})}$ is titrated with standardized 0.300 mol/L $\text{NaOH}_{(\text{aq})}$. What is the pH at the following points:

Reaction:

a) before titration begins

1. ICE Table to determine $[\text{H}_3\text{O}^+]$
2. Approx

3. Solve for x

4. Solve for pH

b) after 10.00 ml of base has been added

1. Calculate moles of acid & base

2. Calculate moles acid remaining (excess)

3. Use moles & T_v to find concentration

4. Use moles of base to find conjugate base formed

5. Concentrate of conjugate base

6. ICE for x

7. Approx

1. Solve for x

9. Solve for pH

c) at the equivalence point

1. What is left in solution at equivalence?
2. Calculate moles of acid & base

3. Calculate the volume required to reach equivalence

4. Calculate the concentration of conjugate base using T_v
5. Calculate K_b & Approx

6. ICE for conjugate base

7. Solve for x

8. Solve for pH

d) select an appropriate indicator to use

e) What is the pH after 35 mL of base has been added?

1. Calculate moles of acid & base

2. Calculate the volume of base required to reach equivalence

3. Calculate extra volume of base

4. Calculate new moles of base at volume left over

5. Calculate the new concentration of base with T_v

6. How many hydroxide ions

7. Solve pOH

8. Solve pH

8.8 BUFFERS

Buffers are mixtures of _____ and _____ or _____ and _____.

These conjugate pairs will allow a solution to resist changes in pH when small amounts of acids and/or bases are added.

Usually buffers have _____ in them that act as a reservoir and help maintain a relatively _____.

Buffer problems should be treated as common ion problems. Le Chatelier's Principle explains the changes.

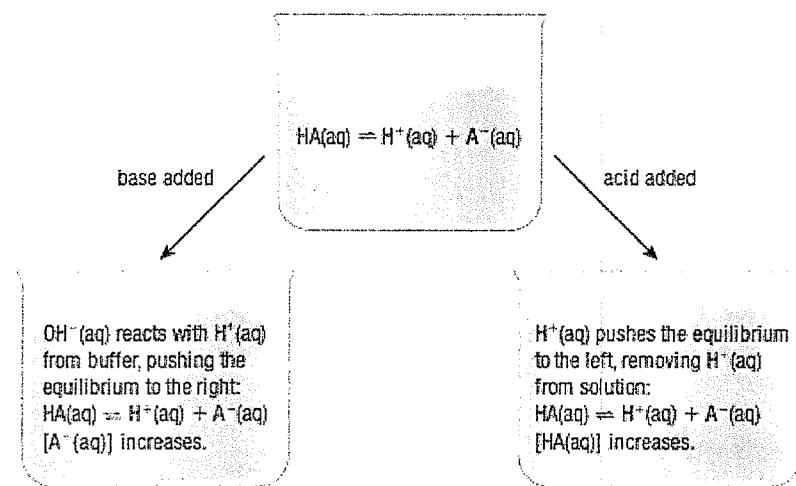


Figure 3 Le Chatelier's principle explains the changes that occur in an acidic buffer solution when a base or an acid is added.

Living organisms are very sensitive to pH changes as enzymes carry out their function optimally over a small pH range. Human blood plasma has a pH of about 7.4 maintained by the hydrogen carbonate- carbonate buffer system. Any change in pH of more than 0.2 induced by poisoning or disease is life-threatening. If blood was not buffered, the acid absorbed from a glass of orange juice would be fatal.

Example Acetic Acid-Acetate ion buffer: $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^-$

When base is added:

When acid is added:

Therefore adding acid or base has _____ since OH^- or H_3O^+ are removed by one component of the buffer solution. Eventually if enough acid or base is added a considerable pH change will occur.

The amount of acid or base that can be added is considered the _____. The buffer capacity is determined by the concentrations of its conjugate acid-base pairs. The highest buffer capacity is when at a point half-way to the equivalence point where $\text{pH} = \text{pK}_a$.

Calculating the pH of a Buffer

Example: Calculate the pH of a buffer that contains 0.20 M acetic acid and 0.20 M sodium acetate.

Henderson-Hasselbalch Equation:

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{conjugate}]}{[\text{acid}]}\right)$$

$$\text{pOH} = \text{pK}_b + \log\left(\frac{[\text{conjugate}]}{[\text{base}]}\right)$$

Ex. 2 Calculate the pH of a buffer containing 0.33 M of ammonia and 0.33 M of ammonium chloride.

Practice Problems Buffers

1. How many grams of sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2$, would have to be added to 1.00 L of 0.15 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, to make a buffer with a pH of 5.00? [22g]

2. How many grams of ammonium chloride would have to be dissolved in 500 mL of 0.20 M NH_3 to prepare a solution buffered at pH 10.00? [0.97g]
3. A buffer solution was prepared by dissolving 2.5 g of ammonium chloride in 125 mL of 0.24 M NH_3 solution. At what pH will this solution serve as a buffer? [9.07]
4. Calculate the pH of a buffered solution made up of 0.015 M sodium acetate and 0.10 M acetic acid. [3.92]

UNIT 3: ELECTROCHEMISTRY

Chapter 10: Electrochemical Cells

Section	Homework
10.1 Galvanic Cells	p. 641 #2 - 6
10.2 Standard Reduction Potentials	p. 647 #1, p. 648 #2 –6, 8a

Learning Goals

1	10.1 Galvanic Cells
I will be able to:	
	<input type="checkbox"/> draw labeled diagrams of galvanic cells showing the oxidation and reduction reactions that occur in each half-cell
	<input type="checkbox"/> label the anode and cathode, the direction of electron flow and ion movement
	<input type="checkbox"/> describe the purpose of the salt bridge in galvanic cells
2	10.2 Standard Reduction Potentials
I will be able to:	
	<input type="checkbox"/> calculate the cell potential using a table of standard reduction potentials
	<input type="checkbox"/> predict the spontaneity of redox reactions based on overall cell potentials

METAL-AIR BATTERIES: PRESENT AND FUTURE

In the future, metal-air batteries could exceed the energy storage abilities of common lithium-ion batteries. Here we look at their potential and compare the different types being developed.

The diagram illustrates a metal-air battery. On the left, a vertical electrode labeled "METAL(M)" has electrons (e-) flowing out of it through an external circuit (represented by a circle with arrows) to the right. On the right, another vertical electrode is shown with arrows pointing towards it from five circles labeled "O2". These O2 molecules are reduced to O2- ions, which then move towards the metal electrode. The entire assembly is enclosed in a rectangular frame.

PERIODIC GRAPHICS

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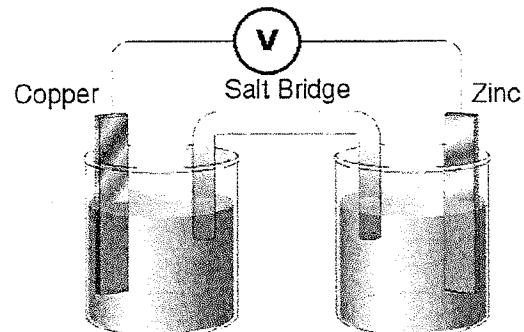
10.1 Galvanic Cells

Galvanic Cells (also called voltaic cells) convert chemical energy into electrical energy, as a result of _____ redox reactions taking place within the cell.

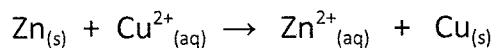
"Electrodes" in Galvanic Cells are typically metal conductors that are submerged in solution.

A cell can be split into two parts connected by a _____ boundary.

Each part, called a "half-cell", consists of one _____ and one _____.



For example, two half-cells might involve copper metal in a solution of copper ions, and zinc metal in a solution of zinc ions.



Electrodes: $\text{Zn}_{(s)}$ and $\text{Cu}_{(s)}$

Electrolytes: $\text{Zn}^{2+}_{(\text{aq})}$ and $\text{Cu}^{2+}_{(\text{aq})}$

Oxidation occurs at the _____, Reduction occurs at the _____

Oxidation: $\text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(\text{aq})} + 2\text{e}^-$ (ANODE)

Reduction: $2\text{e}^- + \text{Cu}^{2+}_{(\text{aq})} \rightarrow \text{Cu}_{(s)}$ (CATHODE)

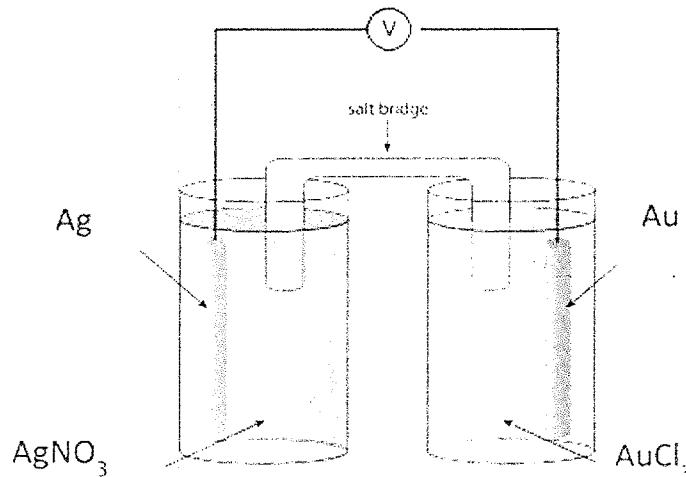
Electrons flow from the anode (oxidation is loss of e^-) to the cathode (reduction is gain of e^-)

Salt Bridge:

- contains an _____ solution that does not interfere in the reaction.
- The ions move out of the salt bridge to maintain electrical _____.
- Cations move toward the cathode and anions toward the anode.
- If we did not have a salt bridge, a positive charge would _____ at the anode and a negative charge would build up at the cathode and electricity would not flow.

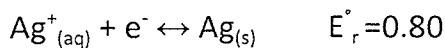
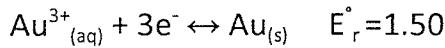
Ex) In the following galvanic cell, which is the anode and the cathode?

Which direction do the electrons flow?



In a _____ cell, the reaction will be _____.

The substance _____ on the list will be the _____, the substance _____ on the list will be the _____.

Standard Reduction Potentials (from table):

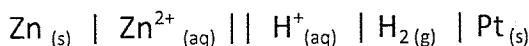
Galvanic Cell Notation:



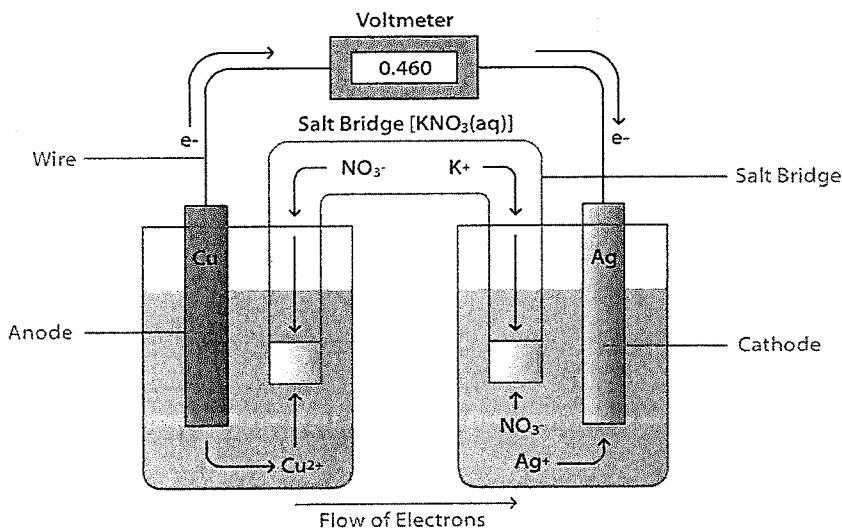
Anode Cathode

- The anode is always shown on the _____ . The cathode is always shown on the _____ .
- | represents a phase boundary between the electrode and the solution in the half cell
- || represents the _____ between 2 half cells
- _____ electrodes are needed if neither substance is a solid.
- Inert electrodes are solid conductors that do not interfere with the desired cell reaction. They provide a location to connect a wire and a surface on which a half-reaction can occur.

E.g., platinum metal foil.

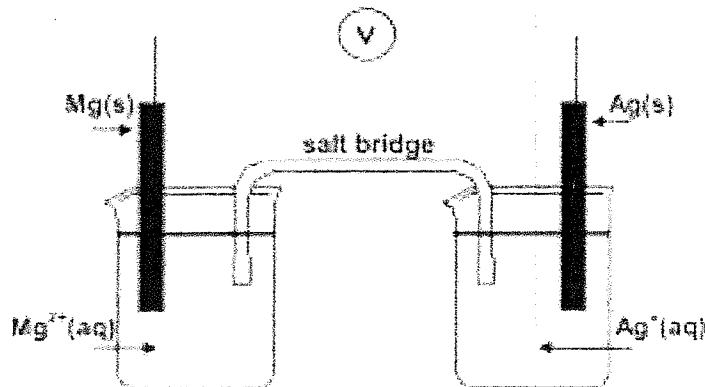


Check your understanding:



- Write the Cell Notation for the Galvanic Cell
- Write the Half Reactions for the Galvanic Cell
- Write the Net Ionic Equation for the Galvanic Cell

4. Consider the following Galvanic Cell:



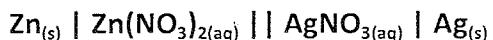
a) Label the direction of electron flow and movement of cations and anions.

b) Write a cell notation

c) Which electrode would decrease in mass? Which electrode would increase in mass?

5. Write the half-reactions and the overall cell reaction.

6. Use the cell notation below to answer the following questions



a) Identify the: Anode _____, Cathode _____ Oxidizing Agent _____, Reducing Agent _____.

b) Write chemical equations to represent the cathode, anode, and overall cell reactions.

c) Draw and label the Galvanic Cell with all components. Use KCl in your salt bridge.

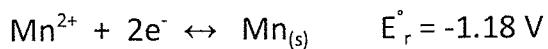
10.2 Calculating Cell Potentials

- The **Cell Potential** is the electric potential difference or voltage between the anode and cathode.
- The **Standard Reduction Potential** is the ability of a half-cell to attract electrons in a cell that is operating under standard conditions (*1.0M, standard pressure and ambient temperature (101.3 kPa, 25°C)*)

Calculating Standard Cell Potentials

If $E^\circ_{\text{(cell)}}$ > 0 then the reaction is spontaneous
If $E^\circ_{\text{(cell)}}$ < 0 then the reaction is not spontaneous

- Ex 1) Determine the cell potential and write the net ionic equation for the following spontaneous reaction:

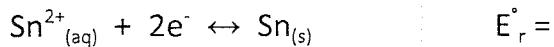
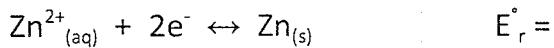


Calculating the Standard Cell Potential:

Net Ionic Equation:

- Since the above reactions are both reduction reactions, you must reverse the reaction occurring at the anode
- Also multiply the equations in order for the electrons to be equal (Note: Do not multiply the reduction potential (E_r°))

Ex 2) Determine the cell potential and write the net ionic equation for the following:



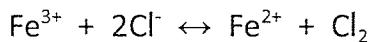
$$E_{(\text{cell})}^\circ = E_{r(\text{cathode})}^\circ - E_{r(\text{anode})}^\circ$$

=

=

Net Ionic Equation:

Ex 3) Determine if the following redox reaction will be spontaneous:



Find the reactions on the standard reduction potential table

Note:

- The above reactions are both reduction reactions
- Since the chlorine is reversed in the net ionic equation, this reaction is occurring at the anode (oxidation)

$$E_{(\text{cell})}^\circ = E_{r(\text{cathode})}^\circ - E_{r(\text{anode})}^\circ$$

$$E_{(\text{cell})}^\circ =$$

Standard Reduction Potentials Table

- A table listing standard reduction potentials of common oxidizing agents and reducing agents in order from strongest to weakest.

Table 1 Standard Reduction Potentials at 25 °C (298 K) for Many Common Half-Reactions

Half-reaction	E°_r (V)	Half-reaction	E°_r (V)
Strongest Oxidizing Agent; Weakest Reducing Agent			
$\text{F}_2(\text{g}) + 2 \text{e}^- \rightleftharpoons 2 \text{F}^-(\text{aq})$	2.87	$\text{Cu}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	0.52
$\text{Ag}^{2+}(\text{aq}) - \text{e}^- \rightleftharpoons \text{Ag}^+(\text{aq})$	1.99	$\text{O}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l}) + 4 \text{e}^- \rightleftharpoons 4 \text{OH}^-(\text{aq})$	0.40
$\text{Co}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Co}^{2+}(\text{aq})$	1.82	$\text{Cu}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	0.34
$\text{H}_2\text{O}_2(\text{aq}) - 2 \text{H}^+(\text{aq}) - 2 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O}(\text{l})$	1.78	$\text{Hg}_2\text{Cl}_2(\text{aq}) + 2 \text{e}^- \rightleftharpoons 2 \text{Hg}(\text{l}) + 2 \text{Cl}^-(\text{aq})$	0.34
$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ce}^{3+}(\text{aq})$	1.70	$\text{AgCl}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) - \text{Cl}^-(\text{aq})$	0.22
$\text{PbO}_2(\text{s}) - 4 \text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{PbSO}_4(\text{s}) - 2 \text{H}_2\text{O}(\text{l})$	1.69	$\text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq}) - \text{H}_2\text{O}(\text{l})$	0.20
$\text{MnO}_4^-(\text{aq}) + 4 \text{H}^+(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{MnO}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$	1.68	$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}^+(\text{aq})$	0.16
$\text{IO}_4^-(\text{aq}) + 2 \text{H}^+(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{IO}_3^-(\text{aq}) - \text{H}_2\text{O}(\text{l})$	1.60	$2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	0
$\text{MnO}_4^-(\text{aq}) + 8 \text{H}^+(\text{aq}) + 5 \text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$	1.51	$\text{Fe}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	-0.036
$\text{Au}^{3+}(\text{aq}) - 3\text{e}^- \rightleftharpoons \text{Au}(\text{s})$	1.50	$\text{Pb}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pb}(\text{s})$	-0.13
$\text{PbO}_2(\text{s}) + 4 \text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Pb}^{2+}(\text{aq}) - 2 \text{H}_2\text{O}(\text{l})$	1.46	$\text{Sn}^{2+}(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	-0.14
$\text{Cl}_2(\text{g}) + 2 \text{e}^- \rightleftharpoons 2 \text{Cl}^-(\text{aq})$	1.36	$\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ni}(\text{s})$	-0.23
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(\text{aq}) - 7 \text{H}_2\text{O}(\text{l})$	1.33	$\text{PbSO}_4(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pb} - \text{SO}_4^{2-}(\text{aq})$	-0.35
$\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O}(\text{l})$	1.23	$\text{Cd}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cd}(\text{s})$	-0.40
$\text{MnO}_2(\text{s}) + 4 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) - 2 \text{H}_2\text{O}(\text{l})$	1.21	$\text{Fe}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	-0.44
$\text{IO}_3^-(\text{aq}) - 6 \text{H}^+(\text{aq}) + 5 \text{e}^- \rightleftharpoons \frac{1}{2} \text{I}_2(\text{s}) + 3 \text{H}_2\text{O}(\text{l})$	1.20	$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cr}^{2+}(\text{aq})$	-0.50
$\text{ClO}_4^-(\text{aq}) - 2 \text{H}^+(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{ClO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$	1.19	$\text{Cr}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	-0.73
$\text{Br}_2(\text{l}) + 2 \text{e}^- \rightleftharpoons 2 \text{Br}^-(\text{aq})$	1.09	$\text{Zn}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
$\text{VO}_2^+(\text{aq}) - 2 \text{H}^+(\text{aq}) - \text{e}^- \rightleftharpoons \text{VO}^{2+}(\text{aq}) - \text{H}_2\text{O}(\text{l})$	1.00	$2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^- \rightleftharpoons \text{H}_2(\text{g}) + 2 \text{OH}^-(\text{aq})$	-0.83
$\text{AuCl}_4^-(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{Au}(\text{s}) + 4 \text{Cl}^-(\text{aq})$	0.99	$\text{Mn}^{2+}(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{Mn}(\text{s})$	-1.18
$\text{NO}_3^-(\text{aq}) + 4 \text{H}^+(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{NO}(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$	0.96	$\text{Al}^{3+}(\text{aq}) - 3 \text{e}^- \rightleftharpoons \text{Al}(\text{s})$	-1.66
$\text{ClO}_2(\text{g}) + \text{e}^- \rightleftharpoons \text{ClO}_2^-(\text{aq})$	0.954	$\text{H}_2(\text{g}) - 2 \text{e}^- \rightleftharpoons 2 \text{H}^+(\text{aq})$	-2.23
$2 \text{Hg}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Hg}_2^{2+}(\text{aq})$	0.91	$\text{Mg}^{2+}(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{Mg}(\text{s})$	-2.37
$\text{Ag}^+(\text{aq}) - \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	0.80	$\text{La}^{3+}(\text{aq}) + 3 \text{e}^- \rightleftharpoons \text{La}(\text{s})$	-2.37
$\text{Hg}_2^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons 2 \text{Hg}^+(\text{aq})$	0.80	$\text{Na}^+(\text{aq}) - \text{e}^- \rightleftharpoons \text{Na}(\text{s})$	-2.71
$\text{Fe}^{3+}(\text{aq}) - \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	0.77	$\text{Ca}^{2+}(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{Ca}(\text{s})$	-2.76
$\text{O}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{O}_2(\text{aq})$	0.68	$\text{Ba}^{2+}(\text{aq}) - 2 \text{e}^- \rightleftharpoons \text{Ba}(\text{s})$	-2.90
$\text{MnO}_4^-(\text{aq}) + \text{e}^- \rightleftharpoons \text{MnO}_4^{2-}(\text{aq})$	0.56	$\text{K}^+(\text{aq}) - \text{e}^- \rightleftharpoons \text{K}(\text{s})$	-2.92
$\text{I}_2(\text{s}) + 2 \text{e}^- \rightleftharpoons 2 \text{I}^-(\text{aq})$	0.54	$\text{Li}^+(\text{aq}) - \text{e}^- \rightleftharpoons \text{Li}(\text{s})$	-3.05
Weakest Oxidizing Agent; Strongest Reducing Agent			

UNIT 4

Chapter 1: Organic Compounds

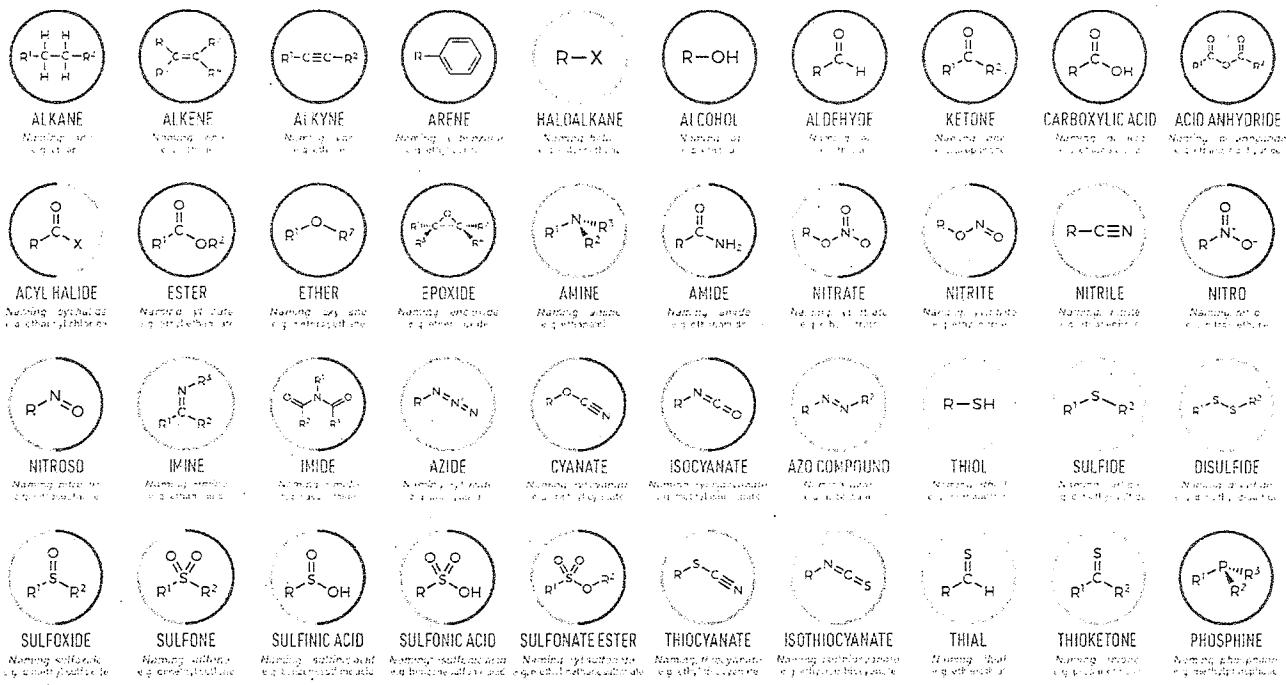
Homework Schedule

Section	Homework	Learning Goals
1.1 Alkanes	p. 14 #1,2 p. 17 #1-5 (2e,4a)	<input type="checkbox"/> Identify parent chains and branches <input type="checkbox"/> Draw and name alkanes
1.2 Alkenes and Alkynes	p. 21 #1,2 (2a) p. 23 #1-3 p. 27 #1-8	<input type="checkbox"/> Draw and name alkenes and alkynes
1.3 Aromatic Hydrocarbons	p. 30 #1,2 p. 31 #1-3, 5, 7	<input type="checkbox"/> Draw and name aromatic compounds
1.4 Alcohols, Ethers, Thiols	p. 34 #1,2 p. 38 #1,2 p. 39 #1,2	<input type="checkbox"/> Draw and name alcohols, ethers and thiols
1.5 Aldehydes & Ketones	p. 41 #1,2 p. 42 #1,2 p. 46 #1,2	<input type="checkbox"/> Draw and name aldehydes and ketones
1.6 Carboxylic Acids, Esters	p. 48 #1,2 p. 50 #1,2 p. 55 #1,2,6	<input type="checkbox"/> Draw and name carboxylic acids and esters
1.7 Amines and Amides	p. 58 #1,2 p. 60 #1,2 p. 62 #1-3	<input type="checkbox"/> Draw and name amines and amides
4.7 Intermolecular Forces	Practice Questions	<input type="checkbox"/> Rank organic molecules based on polarity and strongest IMFs

FUNCTIONAL GROUPS IN ORGANIC CHEMISTRY

Functional groups are the characteristic groups in organic molecules that give them their reactivity. In the formulas below, R represents the rest of the molecule and X represents any halogen atom.

(●) Hydrocarbons (○) Halogen-containing groups (●) Oxygen-containing groups (●) Nitrogen-containing groups (●) Sulfur-containing groups (●) Phosphorus-containing groups



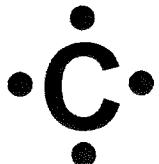
Organic Chemistry: The branch of chemistry concerned with the compounds of carbon; originally confined to compounds produced by living organisms but now extended to include man-made substances based on carbon, such as plastics.

Organic Compounds:	Inorganic Compounds:
<ul style="list-style-type: none"> From living things Contain carbon Over 10 million exist and more than 300,000 are synthesized each year Are formed through covalent bonds 	<ul style="list-style-type: none"> Contain elements other than carbon, although a few do contain carbon A few hundred thousand are known and only a few are synthesized each year Are formed through ionic bonds

Hydrocarbon

- a compound composed only of **carbon** and **hydrogen**

Why are there so many organic compounds? Because carbon is amazing!



- Carbon has 4 valence electrons in the outer shell and can be bound to as many as four other atoms
- Carbon can bond with other carbon atoms effectively to make long chains or rings
- The C-C bond is strong and difficult to break
- Carbon can form strong single, double AND triple bonds

There are 3 types of hydrocarbons:

Formula

Alkane	A hydrocarbon that has only single bonds between carbon atoms	C_nH_{2n+2}
Alkene	A hydrocarbon that contains at least one carbon-carbon double bond	C_nH_{2n}
Alkyne	A hydrocarbon that contains at least one carbon-carbon triple bond	C_nH_{2n-2}

Saturated Hydrocarbons: A hydrocarbon that contains only single bonds. Each carbon atom is bonded to as many H atoms as possible.

Unsaturated Hydrocarbons: A hydrocarbon that contains double/triple bonds, more H atoms can be added.

Structural Diagrams	Example, C_3H_8
Complete structural diagram: shows all the atoms in a structure and the way they are bonded to one another	
Condensed structural diagram: shows the bonds between the carbon atoms but not those between the carbon and hydrogen atoms	
Line structural diagram: shows only lines representing bonds between carbon atoms (it gives us a better sense of 3-D shapes of the molecules)	

Alkane Parent Chain Names

No. of C atoms	Name of alkane	Molecular formula
1	Methane	CH_4
2	Ethane	C_2H_6
3	Propane	C_3H_8
4	Butane	C_4H_{10}
5	Pentane	C_5H_{12}
6	Hexane	C_6H_{14}
7	Heptane	C_7H_{16}
8	Octane	C_8H_{18}
9	Nonane	C_9H_{20}
10	Decane	$\text{C}_{10}\text{H}_{22}$

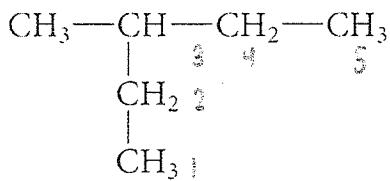
Alkyl Groups

Structure	Name†
—CH ₃	methyl
—CH ₂ CH ₃	ethyl
—CH ₂ CH ₂ CH ₃	propyl
 CH ₃ CHCH ₃	propan-2-yl, or isopropyl
—CH ₂ CH ₂ CH ₂ CH ₃	butyl
 CH ₃ CHCH ₂ CH ₃	butan-2-yl, or sec-butyl
	t-butyl

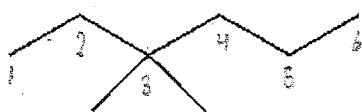
Rules for Naming Alkanes

1. Find the longest carbon chain, called the main chain
 2. Number the carbons such that the branches are on the lowest position #
 3. If there is more than one of the same branch, use a prefix to indicate this (di, tri, tetra, etc)
 4. Write branches in alphabetical order, ignoring any prefixes, including a number to indicate position
 5. Separates #'s and letters with a hyphen, separates #'s with a comma

Example 1:



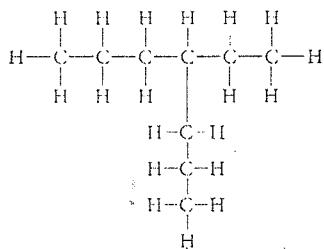
Example 3:



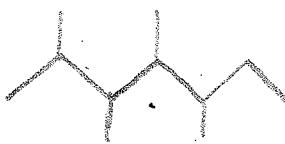
3,3'-dimethylhexane

Example 5: 3,3,4,4-tetramethylheptane

Example 2:



Example 4: 2,3,4,5-tetramethylheptane



Example 6: 4-isopropyl-3,6-dimethyloctane



Structural Isomers

- compounds that have the same molecular formula, but different structures
- different shape and bonding ∴ different properties

Example: C_4H_{10}



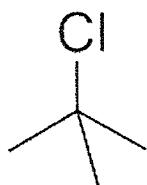
Haloalkanes

- one or more hydrogen atoms have been substituted with halogen atoms (F, Cl, Br, I)

Naming Haloalkanes

- name using a prefix (fluoro-, chloro-, bromo- or iodo-)
- list in alphabetical order and include a position # (alkyl groups have equal priority to halide groups)
- halides are listed alphabetically with alkyl groups

Example 1:



1-chloro-2-methylpropane

Example 2:



1,7-dibromo-1-methylheptane

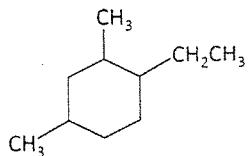
Cycloalkanes

- a regular alkane with a _____ or _____

Naming Cycloalkanes

- when naming, the cyclic structure is always the parent chain, and has the prefix "cyclo"
- number the carbons, in either direction, so that the branches have the lowest possible position number
- when drawing, always use a line diagram for the cyclic chain

Example 1:



Example 2: 1-ethyl-2,4-dimethylcyclopentane

Cyclopropane



Cyclobutane



Cyclopentane



Cyclohexane



Cycloheptane



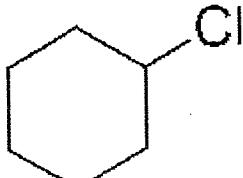
Cyclooctane



Cyclononane



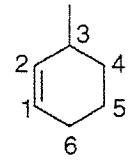
Example 3:



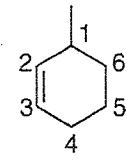
Example 4: 1,1-dimethylcyclobutane

Naming Alkenes and Alkynes

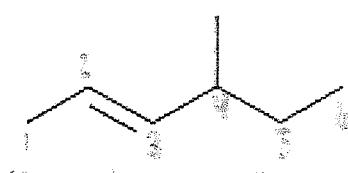
- the parent chain must include double or triple bonds at lowest position number
- alkenes end with -ene, alkynes end with -yne
- position # is shown in front of the ending
- if more than one double bond exists use prefixes (diene, triene, etc)
- when numbering the main chain, double and triple bonds have priority over alkyl groups



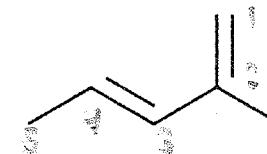
Correct numbering



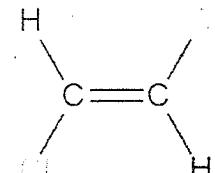
Incorrect numbering

Example 1:

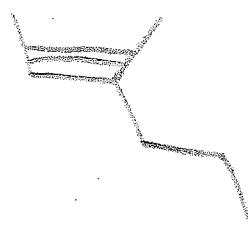
2-methylhex-4-ene

Example 3: 3-methylpent-2-ene**Example 2:**

2-methylpent-1,3-diene

Example 4: 2,5,7-trimethyloct-3-yne**Stereoisomers (Geometric Isomers)**

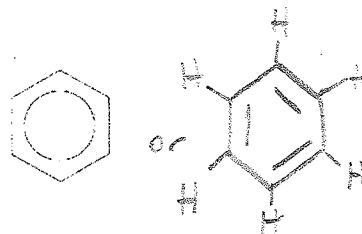
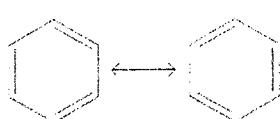
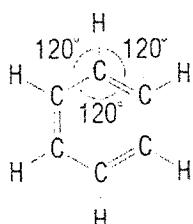
- pairs of molecules which have the same structure but whose functional groups are rotated into a different orientation in three-dimensional space
- cis means that the two groups are on the same side of the double bond
- trans indicates that they are on opposite sides

Example 1: cis-but-2-ene**Example 2: trans-but-2-ene****Example 3: 1,4-dimethylcyclohex-1-ene****Example 3: cis-pent-2-ene****Example 4: trans-pent-2-ene****Example 5: 1,3-dimethylcyclohex-1-ene**

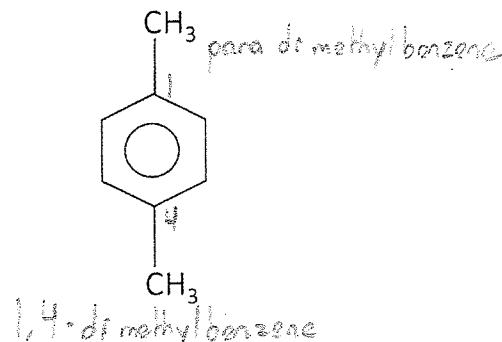
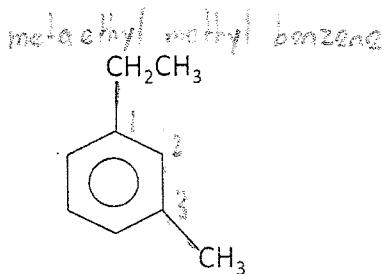
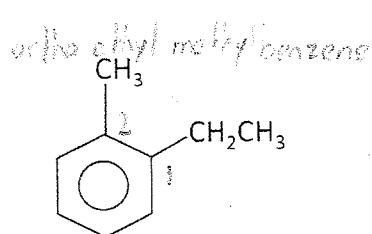
1.3 IUPAC Nomenclature: Aromatics

Aromatic Compounds (includes double and triple bonds)

- an unsaturated hydrocarbon that has a ring structure and a bonding structure that causes it to be chemically stable
- benzene, (C_6H_6), is the simplest aromatic compound



- Electrons are shared over the entire molecule considered 6 identical bonds (halfway between a single and double bond)
- when a benzene ring has only 2 branches, the prefixes **ortho-**, **meta-**, and **para-** are sometimes used instead of numbers



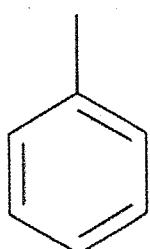
Naming Aromatic Compounds

- if the aromatic compound has non-carbon substituents or an alkyl chain with fewer than 5 carbons
- the name will end in **-benzene**
- if two or more substituents are bonded to a benzene ring, lowest position is assigned alphabetically

Benzene as a Substituent

- if the aromatic compound has a large alkyl chain or an alkyl chain with double or triple bonds then benzene will be a substituent
- the compound is named using phenyl

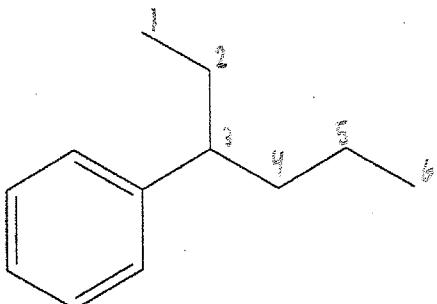
Example 1:



Common name: toluene

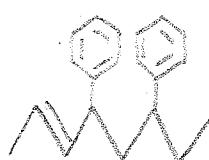
IUPAC: methylbenzene

Example 2:



3-phenylhexane

Example 3: 4,6-diphenyl-oct-2-ene



SUMMARY OF FUNCTIONAL GROUPS

Functional Groups

- A functional group is an element or a group of atoms within a compound that is common to several different substances.
- Compounds with the same functional groups have similar physical and chemical properties. Ex: melting and boiling points, reactivity
- When drawing a carbon chain attached to a functional group we often represent the chain with the letter R to simplify drawings.

Alcohol	Ether	Thiol
Aldehyde	Ketone	Carboxylic Acid
Esters	Amine	Amide

Alcohols

- an alcohol is an organic compound that contains the _____ functional group (_____)
- an alcohol can be primary (1°), secondary (2°) or tertiary (3°)

Primary (1°)

-OH is bonded to C
bonded to only 1 other C

Secondary (2°)

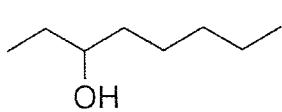
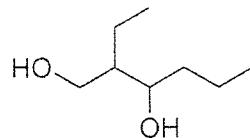
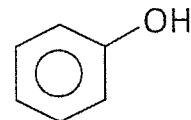
-OH is bonded to a C
bonded to 2 other C

Tertiary (3°)

-OH is bonded to C
bonded to 3 other C

Naming an Alcohol

1. find the longest chain that contains the -OH group
2. name the parent chain, replace the -e ending with -ol
3. for more than one - OH group, use prefixes (diol, triol, etc)
4. number the parent chain to give priority to -OH
5. include the position # of -OH in front of -ol ending
6. number and name any alkyl substituents
7. if another functional group has priority over the alcohol, the alcohol substituent is named *-hydroxy*

Example 1:**Example 2:****Example 3:****Example 4:** 3-cyclopropylpentan-2-ol**Example 5:** pentane-1,5-diol**Properties of Alcohols**

- Because of the hydroxyl (OH) group, they can form hydrogen bonds with other molecules which leads to **higher boiling points and greater solubility** in water than alkanes of the same size

Methane	Chloromethane	Methanol
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Example 1: Draw the following structures, list the strongest IMF and rank from lowest to highest boiling point.

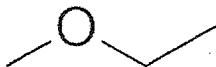
Ethers

- an ether has two alkyl groups joined by an oxygen atom, ethers can be used as anaesthetics

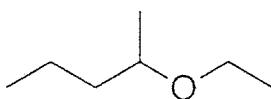
Naming an Ether

- name the longest alkane chain as the parent chain
- name the alkyl substituent attached to the oxygen, change the ending from *-yl* to *-oxy*
- include the position # in front showing where the oxygen is attached to the parent chain
- number and name any alkyl substituents
- oxy* substituents are listed before alkyl substituents

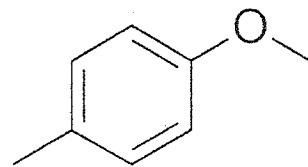
Example 1:



Example 2:



Example 3:



Example 4: 1-methoxypropane

Example 5: 2-methoxypropane

Example 6: 2-methoxy-3-ethylpent-4-ene

Properties of Ethers

- Like alkanes, ethers are very unreactive substances. Pure ethers cannot form hydrogen bonds. Ethers are used as anaesthetics. Have lower boiling points than similar alcohols
- Slightly higher boiling points than hydrocarbons of the same length because the C-O bond is still a polar bond, so ethers are more polar than regular hydrocarbons

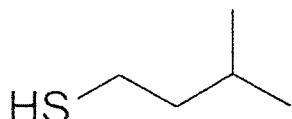
Thiols R-SH

- contain the sulfhydryl functional group
- name by adding ending "thiol"
- generally have a strong odour (ex. skunk, garlic, sewage, additive to natural gas)

Naming a Thiol

- Locate the longest chain that contains the -SH group and name it
- Name the alkane and add the ending *-thiol* (for more than 1 -SH, "dithiol", "trithiol")
- Add a position # to show the position of -SH (on lowest carbon)
- Number and name any branches

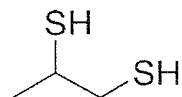
Example 1:



Example 2:

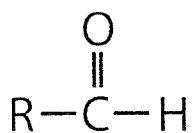


Example 3:

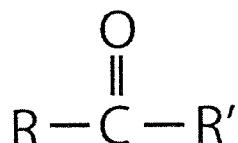


Aldehyde

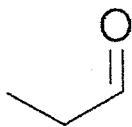
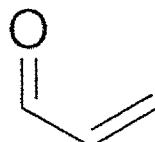
- has the C=O at the **end** of a parent chain

**Ketone**

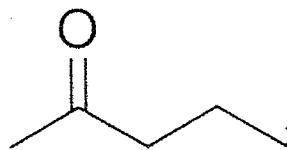
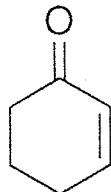
- has an C=O **within** the parent chain

**Naming Aldehydes**

- start numbering the chain so that the aldehyde group is at position #1
- replace with the "e" ending with *-al*
- number and name any alkyl substituents
- condensed structures show aldehydes as R-CHO

Example 1:**Example 2:****Example 3:** 5-ethyl-4,6-dimethyloctanal**Naming Ketones**

- replace with the ending *-one*
- number so that the ketone is at the lowest possible position
- indicate the ketone position in front of the -one
- number and name any alkyl substituents
- condensed structures show ketone as R-C(O)-R'

Example 1:**Example 2:****Example 4:** 5-bromoheptan-3-one**Example 5:** 4-hydroxypentan-2-one**Properties** Draw the following structures, list the strongest IMF and rank from lowest to highest solubility

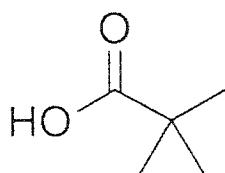
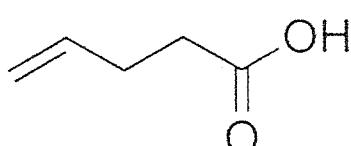
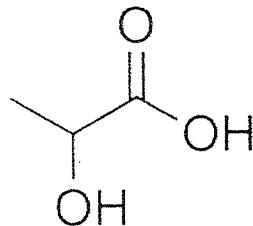
Ethane	Ethanol	Ethanal
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Carboxylic Acids

- contain the carboxyl group located at the end of a parent chain

Naming Carboxylic Acids

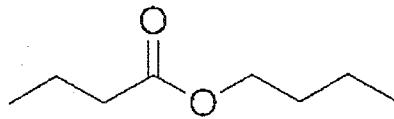
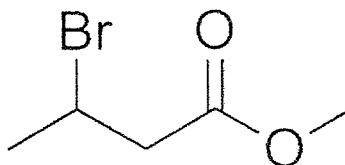
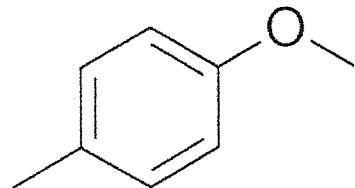
- position #1 will be where the carboxyl group is
- replace the ending -e with -oic acid
- number and name any alkyl substituents

Example 1:**Example 2:****Example 3:****Esters**

- contain the carboxyl group between two carbon chains

Naming an Ester

- Name the parent chain, which includes the C=O group, replace the -e ending with -oate
- Name the alkyl group attached to -O-, indicate position of attachment to the alkyl group (if not #1)
- Put the 2 names together (but as separate words)

Example 1:**Example 2:****Example 3:****Properties of Esters**

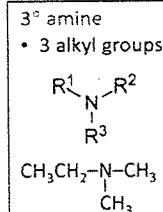
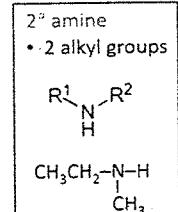
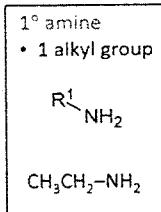
- Esters have a carbonyl group (C=O), but no carboxyl group (C=O and OH). This makes them **less soluble** in water and they have **lower boiling points** than their corresponding carboxylic acids
- Most esters have a sweet, fruity odour (found in perfumes, candy, fruit)

Properties Draw the following structures, list the strongest IMF and rank from lowest to highest solubility

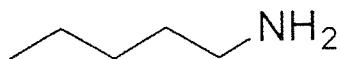
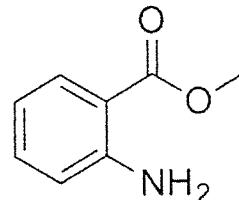
butanoic acid	methyl propanoate	Butan-2-one
---------------	-------------------	-------------

Amines:

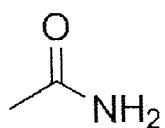
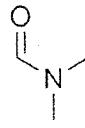
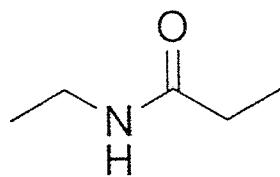
- An organic compound derived from ammonia (NH_3) by replacement of one or more hydrogen atoms with alkyl groups.
- Classified according to the # of alkyl groups attached to the N

**Naming Amines**

- Identify the longest chain attached to the N atom as the parent chain.
- Replace the –e at the end with the name –amine.
- Indicate position #.
- Name other alkyl groups using N- instead of a position # if attached to the nitrogen
- List alkyl groups attached to N before groups attached to parent chain.

Example 1:**Example 2:****Example 3:****Amides:** contains a carbonyl group bonded to a nitrogen atom**Naming Amides**

- The parent chain always includes the C=O. This carbon is position #1.
- Replace –e ending with –amide
- Name any alkyl groups attached to the N, using N-, and those attached to the parent chain using position #

Example 1:**Example 2:****Example 3:****Properties of Amines & Amides**

- The N-C and N-H bonds are polar and they can form hydrogen bonds, this means that
- amine have **higher** boiling and melting points than similar-sized hydrocarbons
- smaller amines are soluble in water act as weak bases
- Some smaller amides are soluble due to polar -NH groups

Properties Draw the following structures, list the strongest IMF and rank from lowest to highest solubility

methanamide

methane

methanol

Intermolecular Forces Review

	Electrostatic Attractions	Covalent Bonds	Intermolecular Forces (van der Waals forces) in Covalent Compounds			
			in Ionic Compounds	in Molecular Compounds	Hydrogen Bonding	Ion Dipole Forces
Example	NaCl	CH_4	B/w water molecules	$\text{Na}^+ & \text{H}_2\text{O}$	$\text{H}-\text{O}-\text{H}$	$\text{Na}^+ \cdots \text{O}^-$
Diagram						
When does this force occur?	Positive ion attracted to a negative ion	Elements share electrons	B/W compounds that have H and N, O or F	B/W ions and polar molecules	B/w polar molecules (not with H/N,O,F)	B/w nonpolar molecules
Relative Strength compared to other forces	Strongest	Weakest	2 nd Strongest Van der Waals	Strongest Van der Waals	3 rd Strongest Van der Waals	Weakest Van der Waals
What determines strength of force?			large difference in electronegativity between the H atom and the N, O or F		Difference in electronegativity between atoms in molecule	Temporary Dipole created in atoms

Watch: Intermolecular Forces in Organic Molecules

Bigger molecules can have a large dispersion interaction because of their molecular weight.

Examples: CH_4 C_5H_{12} CCl_4

Molecular Weight: _____

Boiling Point: _____

Explanation:

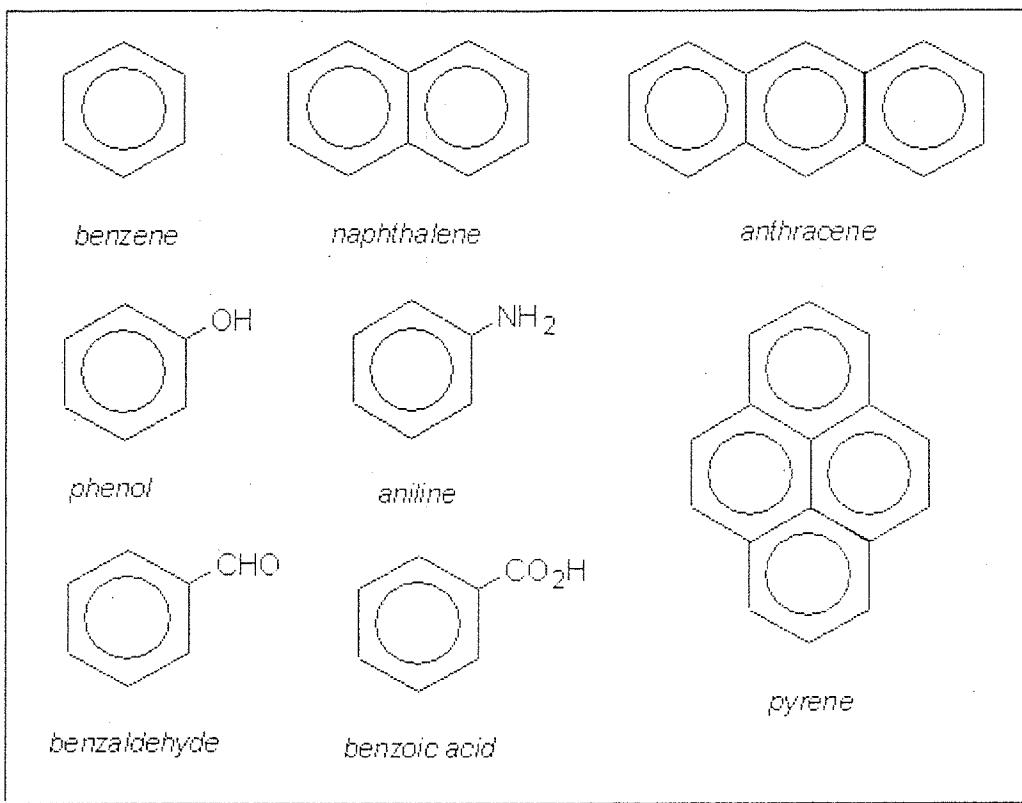
Boiling point increases with molecular mass because of their stronger intermolecular forces.

	Cyclopentane	THF	Butanal	Butanol
MW (g/mol)				
LDF				
Dipole-Dipole				
Hydrogen Bonding				
Boiling Point (degrees Celsius)				

Organic Compounds listed in order of increasing solubility in water AND melting/boiling points:

↑	1° Amide	Because primary amides have two NH bonds, they have even stronger hydrogen bonds than carboxylic acids.
	2° and 3° Amides	The CN and NH bonds are polar. Secondary amides also experience hydrogen bonding. Their solubility decreases as the non-polar hydrocarbon part of the molecule increases in size.
	Carboxylic Acids	The OH and CO bonds make carboxylic acids polar. Hydrogen bonding occurs between carboxylic acid molecules as well as water molecules. The solubility of carboxylic acids in water decreases as the number of carbon atoms increases.
	Amines	CN and NH bonds are polar. The presence of one or more NH bonds allows hydrogen bonding to take place. The solubility of an amine decreases as the number of carbon atoms increases. The boiling points of primary and secondary amines (which contain NH bonds) are higher than the boiling points of tertiary amines (which do not contain an NH bond).
	Alcohols	The OH bond is very polar. Alcohols experience hydrogen bonding with other alcohol molecules and with water. The solubility of an alcohol decreases as the number of carbon atoms increases.
	Alkyl Halides	Alkyl halides are similar in structure, polarity, and reactivity to alcohols.
	Ester	Esters are polar. They contain a CO bond. Therefore, esters can accept hydrogen bonds from water but not from other ester molecules.
	Aldehydes and Ketones	The CO bond is polar, so aldehydes and ketones are polar. They can accept hydrogen bonds from water but not from other aldehyde or ketone molecules.
	Ether	The CO bond makes ethers polar. Hydrogen bonding cannot occur between ether molecules but can occur between an ether molecule and water.
	Hydrocarbons	Hydrocarbons – non polar – the longer they are, the more dispersion forces they have, and the higher the melting/boiling point. But they are NOT soluble in water.

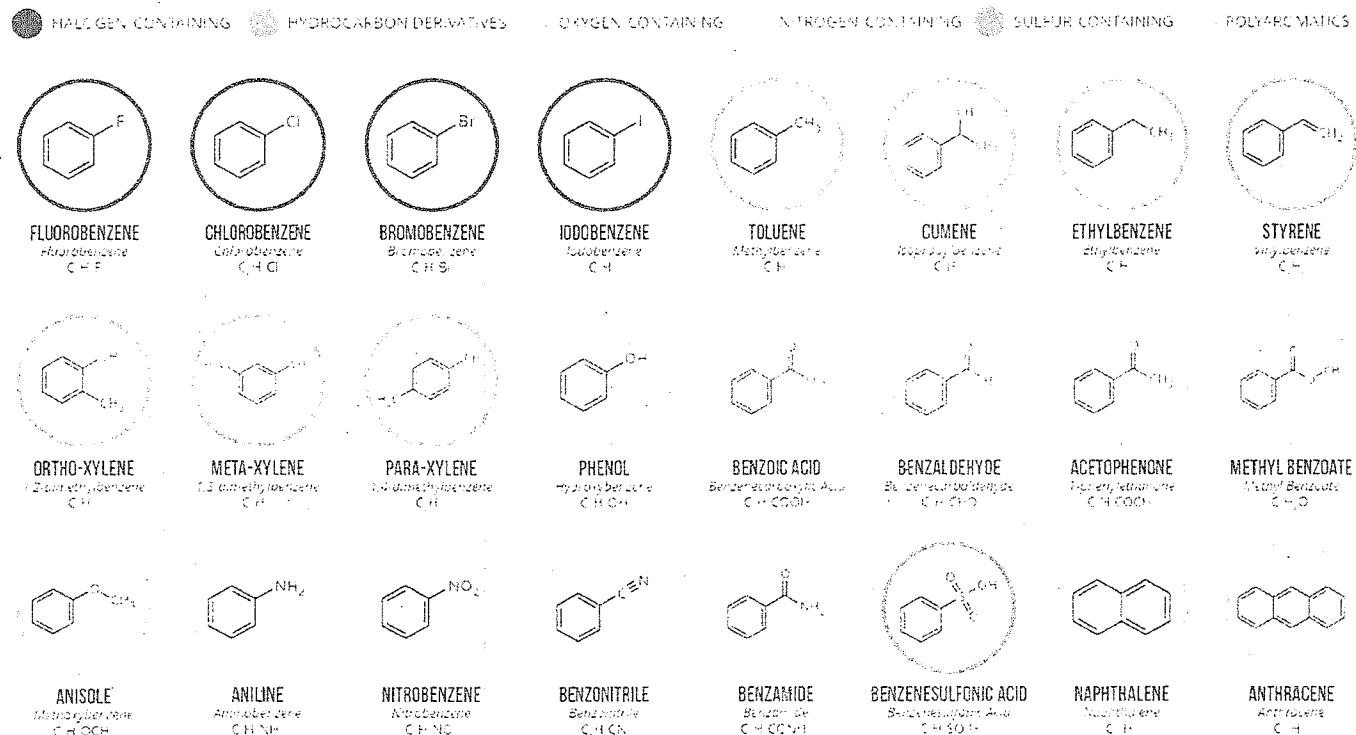
Common Benzene Derivatives To Remember...



toluene

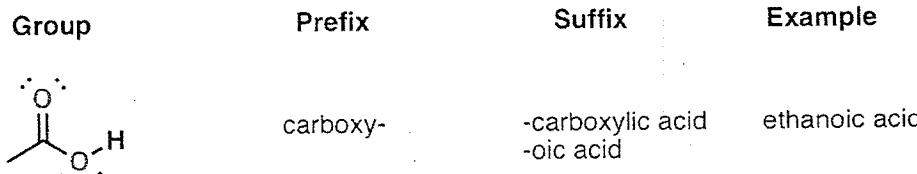
BENZENE DERIVATIVES IN ORGANIC CHEMISTRY

A WIDE VARIETY OF IMPORTANT ORGANIC COMPOUNDS ARE DERIVED FROM BENZENE, BY REPLACING ONE OF THE HYDROGENS WITH A DIFFERENT FUNCTIONAL GROUP. THEY CAN HAVE BOTH COMMON & SYSTEMATIC NAMES, WHICH CAN BE CONFUSING. COMMON NAMES ARE SHOWN IN BOLD, WITH SYSTEMATIC NAMES SHOWN IN ITALICS.

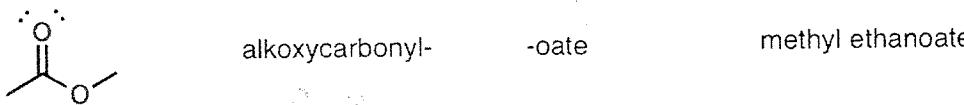


1,2-Dimethylpropano-2-one 3-methoxybenzoic acid 3-Isobutyl-3-methylcyclohexanone

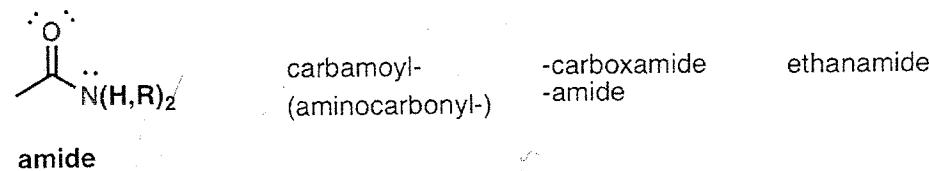
Organic Functional Group Priority List...



carboxylic acid



ester



amide



aldehyde



ketone



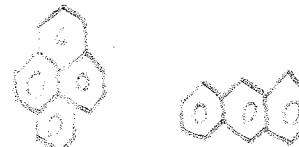
alcohol



thiol



amine



3-methoxybenzoic acid 3-isobutyl-3-methylcyclohexanone

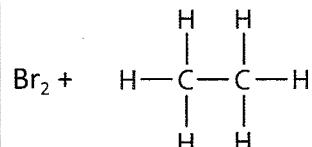
Organic Chemistry Reactions Review

For each of the following reactions, 1) name the reactants, 2) draw and name the products, 3) be specific as possible in the type of reaction.

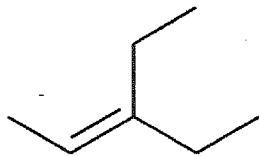
Type of reaction: _____



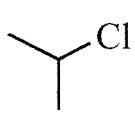
Type of reaction: _____



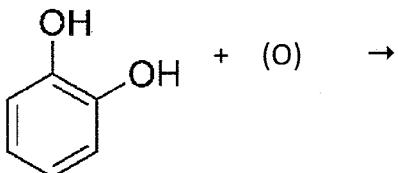
Type of reaction: _____



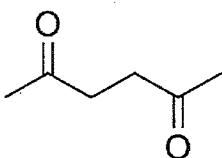
Type of reaction: _____



Type of reaction: _____

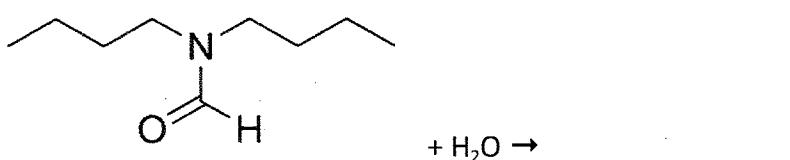


Type of reaction: _____

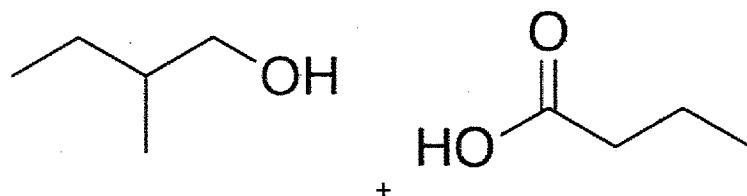


Organic Chemistry Reactions Review

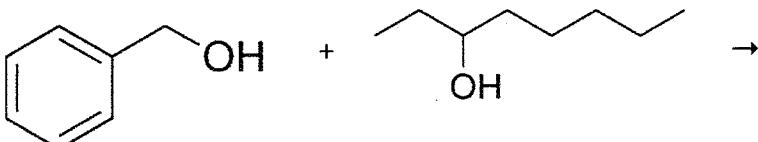
Type of reaction: _____



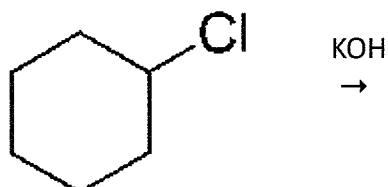
Type of reaction: _____



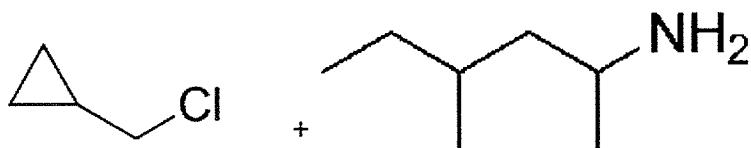
Type of reaction: _____



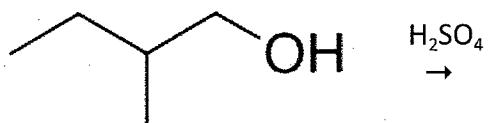
Type of reaction: _____



Type of reaction: _____



Type of reaction: _____

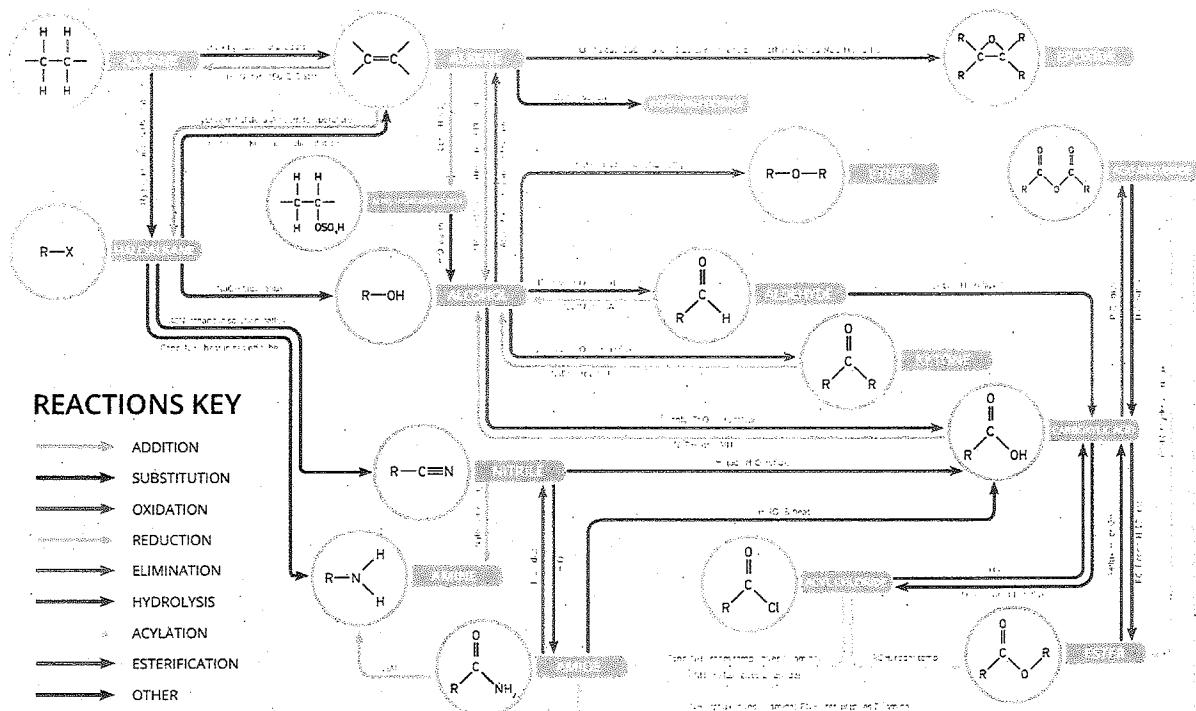


Chapter 1 & 2 Organic Reactions & Polymers

Homework Schedule

Section	Homework
1.2 Alkenes and Alkynes	p. 26 #1,2 p. 27 #9,10
1.3 Aromatic Hydrocarbons	p. 31 #3,4,6
1.4 Alcohols, Ethers, Thiols	p. 37 #1, 2 p. 39 #3 – 6
1.5 Aldehydes and Ketones	p. 45 # 1 – 3 p. 46 #3 – 5
1.6 Carboxylic Acids, Esters, and Fats	p. 53 #1 p. 55 #3,4,8,9
1.7 Amines and Amides	p. 62# 4,5,6
2.1 Introducing Polymers	p. 83# 1 – 4
2.2 Synthetic Addition Polymers	p. 87 # 1 – 3, p. 93 # 1- 3, 7
2.4 Synthetic Condensation Polymers	p. 98 #1, p. 99 #1 – 3
2.6 Natural Polymers	p. 105# 3,4

ORGANIC REACTIONS MAP



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Reactions of Alkanes

1. Combustion
2. Substitution

Combustion: Reaction with oxygen

Example: Combustion of Propane

Substitution: A hydrogen atom is being substituted for a halogen (Group 17)

Example: Substitution of Methane

Reactions of Alkenes & Alkynes

1. Combustion
2. Addition

Halogenation	Hydrohalogenation	Hydration	Hydrogenation
--------------	-------------------	-----------	---------------

Halogenation (add halogens) Example:**Hydrogenation (add hydrogen gas) Example:****Hydrohalogenation (add hydrogen and halogens) Example:****Hydration (add H^+ and OH^-) Example:**

Markvnikov's Rule "Rich get richer"

- The carbon within the double bond that already has more hydrogens will bond to the H from H_2O & HX
- Carbon that has more hydrogens gets "richer" with hydrogens

Example:

Reaction of Aromatic Hydrocarbons: Only undergo **substitution** reactions. Example:

Nitration:

Alkylation:

Reactions of Haloalkanes

1. Substitution (Alkanes \rightarrow Haloalkane)
2. Addition (Alkenes \rightarrow Haloalkane)
3. Elimination (Haloalkane \rightarrow Alkene)

Elimination Example:

Reactions of Alcohol

1. Combustion

2. Oxidation

3. Dehydration

Dehydration of Butan-2-ol (_____)

MAJOR

MINOR

Oxidation

Example 1: Oxidation of a _____ alcohol)

Example 2: Oxidation of a _____ alcohol)

Example 3 : Oxidation of a _____ alcohol)

Reactions of Ethers

1. Condensation

2. Hydrolysis

Condensation: Formation of an Ether

Hydrolysis: Breaking ether bond

Reactions of Aldehydes

1. Hydrogenation

2. Oxidation

Hydrogenation Example:

Oxidation Example:

Reactions of Ketones

1. Hydrogenation

2. Oxidation

Hydrogenation

Example:

Oxidation Example:

Reactions of Carboxylic Acids

1. Esterification

2. Hydrolysis

Esterification

Hydrolysis Example:

Reactions of Amines

1. Substitution

2. Condensation

Substitution:

Example 1: Substitution with ammonia

Example 2: Substitution with a primary amine

Condensation:

Example 1: Carboxylic Acid and Ammonia

Example 2: Carboxylic Acid and primary amine

Reactions of Amides

Hydrolysis: Example

SCH 4U

Polymers

A polymer is a very long molecule that is made by linking together many smaller molecules called monomers

There are two types of polymers:

1) **Synthesis:** man-made polymers

Eg: plastics

2) **Natural:** polymers made by nature

Eg: cellulose, starch, proteins

The name of a polymer is usually written with the prefix **poly** (meaning many) before the name of the monomer. Often the common name is used for the monomer rather than the IUPAC name.

There are two ways to form synthetic polymers:

- **Addition** Polymerization
- **Condensation** Polymerization

Addition Polymerization:

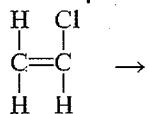
- A very long organic molecule formed as a result of addition reactions b/w carbon atoms with double bonds
- Monomers with double bonds are joined together through multiple addition reactions to form a polymer

Example 1:

Condensed form of a polymer:

Name of polymer:

Example 2: For the following monomer, draw three repeating units of the polymer in condensed form.



Polyvinyl chloride (PVC)

Condensation Polymerization:

- A very long organic molecule formed as a result of condensation reactions between monomers with two functional groups
- Monomers are joined together by the formation of **ester** or **amide** bonds; a second smaller product (usually **water**) is produced
- Each monomer **MUST** have two functional groups (usually one at each end of the molecule)

Types of Condensation Polymers

- A. Condensation polymers that contain **ester** bonds are called **polyesters**. They form from the esterification of diacids and dialcohols.

Example:

- B. Condensation polymers that contain **amide** bonds are called **nylon** or **polyamides**. They are formed from dicarboxylic acids and diamines.

Example:

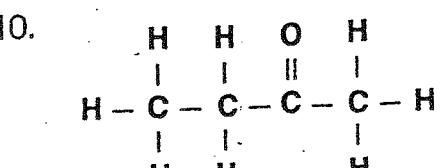
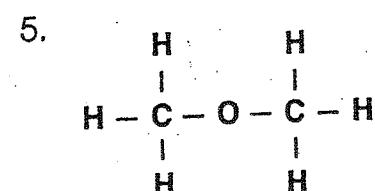
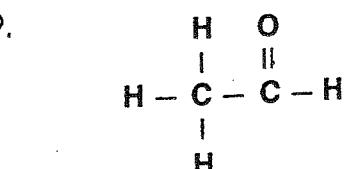
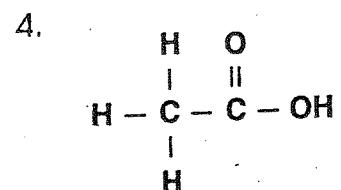
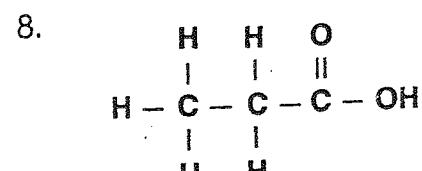
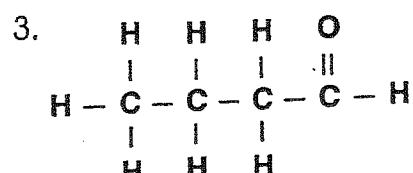
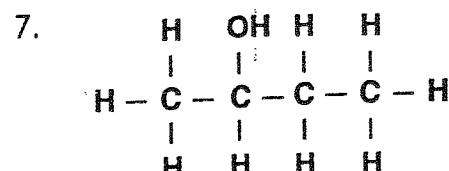
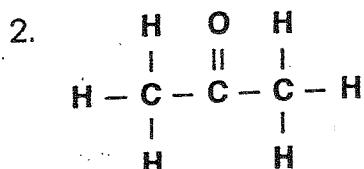
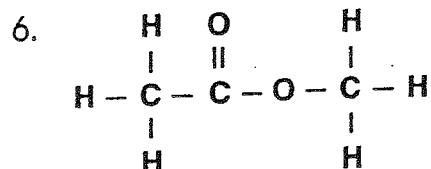
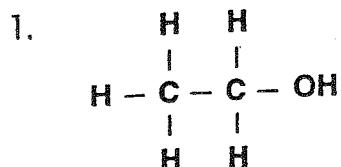
Polymer Cross-Linking

Chemical bonds can form between separate polymer strands in a process called cross linking. Cross-links in polymers make the resulting plastics stronger and more rigid.

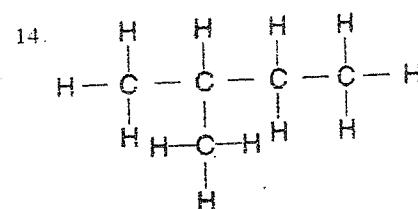
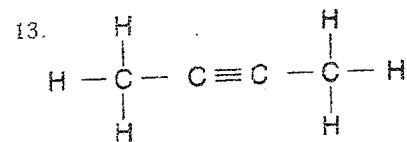
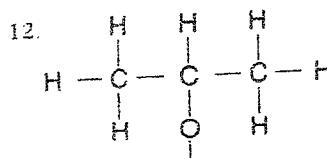
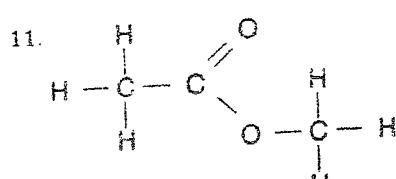
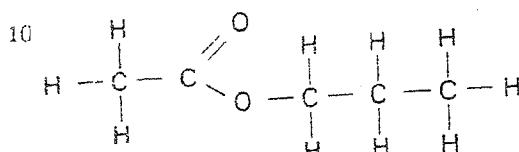
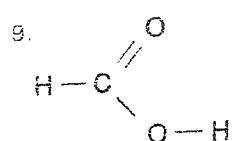
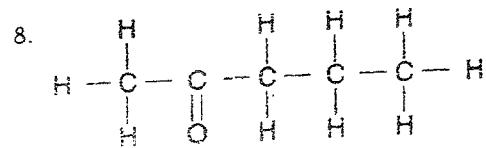
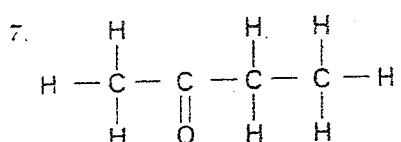
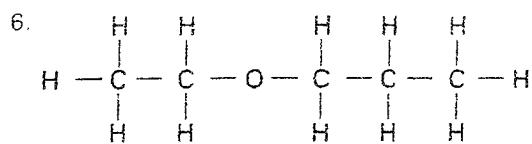
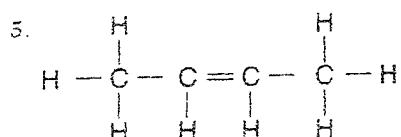
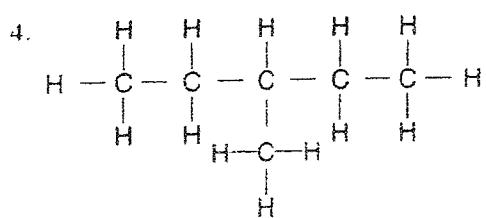
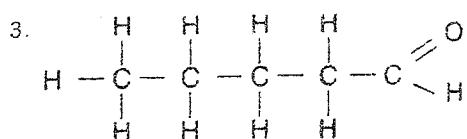
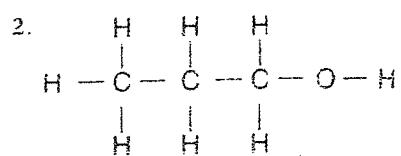
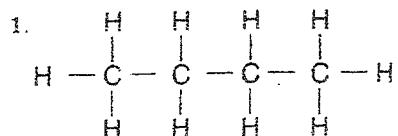
Example: Divinylbenzene

SCH 4C**FUNCTIONAL GROUPS**

Classify each of the following molecules.

SET #1

SET #2



UNIT 5 STRUCTURE & PROPERTIES OF MATTER

CHAPTER 3: ATOMS

Section	Homework
3.4 Quantum Numbers	p. 159 #6 – 9
3.5 Atomic Structure and the Periodic Table	p. 166 # 1 – 4, p. 170 #1 – 2, p. 172 #4, 5, 6, 9

A HISTORY OF THE ATOM: THEORIES AND MODELS

How have our ideas about atoms changed over the years? This graphic looks at atomic models and how they developed.

The diagram illustrates the progression of atomic models through five stages:

- SOLID SPHERE MODEL (1803):** Shows a uniform sphere with a central positive charge.
- PLUM PUDDING MODEL (1904):** Shows a sphere containing numerous small negative charges (plums) embedded in a positive background (pudding).
- NUCLEAR MODEL (1911):** Shows a central nucleus with electrons as small dots scattered around it.
- PLANETARY MODEL (1913):** Shows electrons orbiting the central nucleus in elliptical paths.
- QUANTUM MODEL (1926):** Shows electrons in discrete orbits around the nucleus, represented by clouds of probability.

JOHN DALTON (1803)

Dalton drew upon the Ancient Greek idea of atoms (the word 'atom' comes from the Greek 'atomos' meaning indivisible). His theory stated that atoms are indivisible, those of a given element are identical, and compounds are combinations of different types of atoms.

J.J. THOMSON (1904)

Thomson discovered electrons (which he called 'corpuscles') in atoms in 1897, for which he won a Nobel Prize. He subsequently 'produced' the 'plum pudding' model of the atom. It shows the atom as composed of electrons scattered throughout a spherical cloud of positive charge.

ERNEST RUTHERFORD (1911)

Rutherford fired positively charged alpha particles at a thin sheet of gold foil. Most passed through with little deflection, but some deflected at large angles. This was only possible if the atom was mostly empty space, with the positive charge concentrated in the centre: the nucleus.

NIELS BOHR (1913)

Bohr modified Rutherford's model of the atom by stating that electrons moved around the nucleus in orbits of fixed sizes and energies. Electron energy in this model was quantised; electrons could not occupy values of energy between the fixed energy levels.

ERWIN SCHRÖDINGER (1926)

Schrödinger stated that electrons do not move in set paths around the nucleus, but in waves. It is impossible to know the exact location of the electrons; instead, we have 'clouds of probability' called orbitals, in which we are more likely to find an electron.

RECOGNISED ATOMS OF A PARTICULAR ELEMENT DIFFER FROM OTHER ELEMENTS

ATOMS AREN'T INVISIBLE - THEY'RE COMPOSED FROM SUBATOMIC PARTICLES

RECOGNISED ELECTRONS AS COMPONENTS OF ATOMS

NO NUCLEUS! DIDN'T EXPLAIN LATER EXPERIMENTAL OBSERVATIONS

REALISED POSITIVE CHARGE WAS LOCALISED IN THE NUCLEUS OF AN ATOM

DID NOT EXPLAIN WHY ELECTRONS REMAIN IN ORBIT AROUND THE NUCLEUS

PROPOSED STABLE ELECTRON ORBITS, EXPLAINED THE EMISSION SPECTRA OF SOME ELEMENTS

MOVING ELECTRONS SHOULD EMIT ENERGY AND COLLAPSE INTO THE NUCLEUS. MODEL DID NOT WORK WELL FOR HEAVIER ATOMS

SHOWS ELECTRONS DON'T MOVE AROUND THE NUCLEUS IN ORBITS, BUT IN CLOUDS WHERE THEIR POSITION IS UNCERTAIN

STILL WIDELY ACCEPTED AS THE MOST ACCURATE MODEL OF THE ATOM

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3.4 QUANTUM NUMBERS

Schrodinger's wave equation describes the **quantized energies of the electron** in an atom and the wave function that determines **the probability of finding electrons** in various regions of the atom.

Solving the wave equation gives a set of _____, and their corresponding _____. Electrons are found in these orbitals, which have various 3D shapes and sizes.

Quantum numbers are like addresses for locating the **position of an electron** by its city street, number and apartment number.

An orbital is described by a set of four quantum numbers:

- Principal quantum number, _____ : _____
- Angular momentum (secondary) quantum number, _____ : _____
- Magnetic quantum number, _____ : _____
- Spin quantum number, _____ : _____

Principal Quantum Number, n

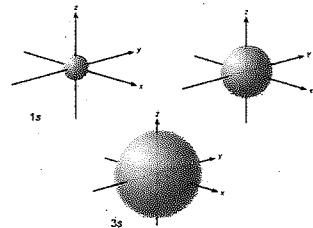
- indicates the relative _____ and _____ of atomic orbitals
- represents the energy level, or shells, the electrons can occupy in an atom
- $n = 1, 2, 3, 4 \dots$
- as n _____, orbital size _____
- maximum number of electrons per energy level: _____

Angular Momentum (Secondary) Quantum Number, ℓ

- energy sublevels are contained within the principal energy levels
- $\ell = \text{_____}$ each energy sublevel (s, p, d , or f) relates to orbitals of different shape

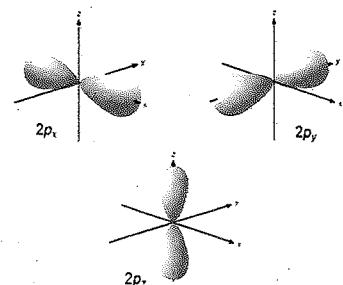
1. s orbitals

- $\ell = \text{_____}$
- each energy level has _____ s orbital
- s orbitals can hold _____



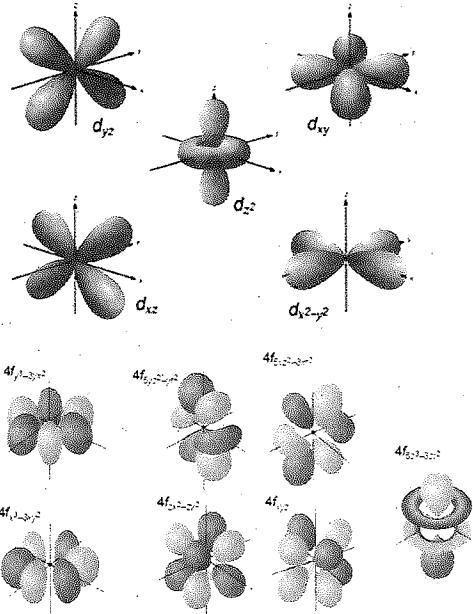
2. p orbitals

- $\ell = \text{_____}$
- p orbitals have greater energy than s orbitals
- _____ p orbitals per subshell
- each p orbital can hold _____ for a total _____ in the p subshell
- _____



3. *d* orbitals

- $\ell = \underline{\hspace{2cm}}$
- *d* orbitals have greater energy than *p* orbitals
- $\underline{\hspace{2cm}}$ *d* orbitals in each subshell
- each *d* orbital can hold $\underline{\hspace{2cm}}$, for a total $\underline{\hspace{2cm}}$ in the *d* subshell



4. *f* Orbitals

- $\ell = \underline{\hspace{2cm}}$
- *f* orbitals have greater energy than *d* orbitals
- $\underline{\hspace{2cm}}$ *f* orbitals in each subshell
- each *f* orbital can hold $\underline{\hspace{2cm}}$, for a total $\underline{\hspace{2cm}}$ in the *f* subshell

Magnetic Quantum Number, m_ℓ

- represents the orientation of the subshells
- has values of $\underline{\hspace{2cm}}$ to $\underline{\hspace{2cm}}$

$$s = \begin{array}{|c|} \hline \quad \\ \hline \end{array} \quad 0$$

$$p = \begin{array}{|c|c|c|} \hline & & \\ \hline -1 & 0 & 1 \\ \hline \end{array}$$

$$d = \begin{array}{|c|c|c|c|c|} \hline & & & & \\ \hline -2 & -1 & 0 & 1 & 2 \\ \hline \end{array}$$

$$f = \begin{array}{|c|c|c|c|c|c|c|c|} \hline & & & & & & & \\ \hline -3 & -2 & -1 & 0 & 1 & 2 & 3 \\ \hline \end{array}$$

Spin Quantum Number, m_s

- represents the spin of the electron
- has values of $\underline{\hspace{2cm}}$ or $\underline{\hspace{2cm}}$

Example 1: For a given atom, $n = 3$.

(a) What are all possible values of ℓ for this atom?

(b) What are all possible values for m_ℓ for this atom?

(c) What is the maximum number of electrons this atom could have in this energy level?

Example 2: Which of the following sets of Quantum Numbers could NOT exist?

(a) $n = 3 \quad \ell = 2 \quad m_\ell = -1 \quad m_s = +\frac{1}{2}$

(b) $n = 2 \quad \ell = 2 \quad m_\ell = 0 \quad m_s = -\frac{1}{2}$

(c) $n = 3 \quad \ell = 1 \quad m_\ell = -2 \quad m_s = +\frac{1}{2}$

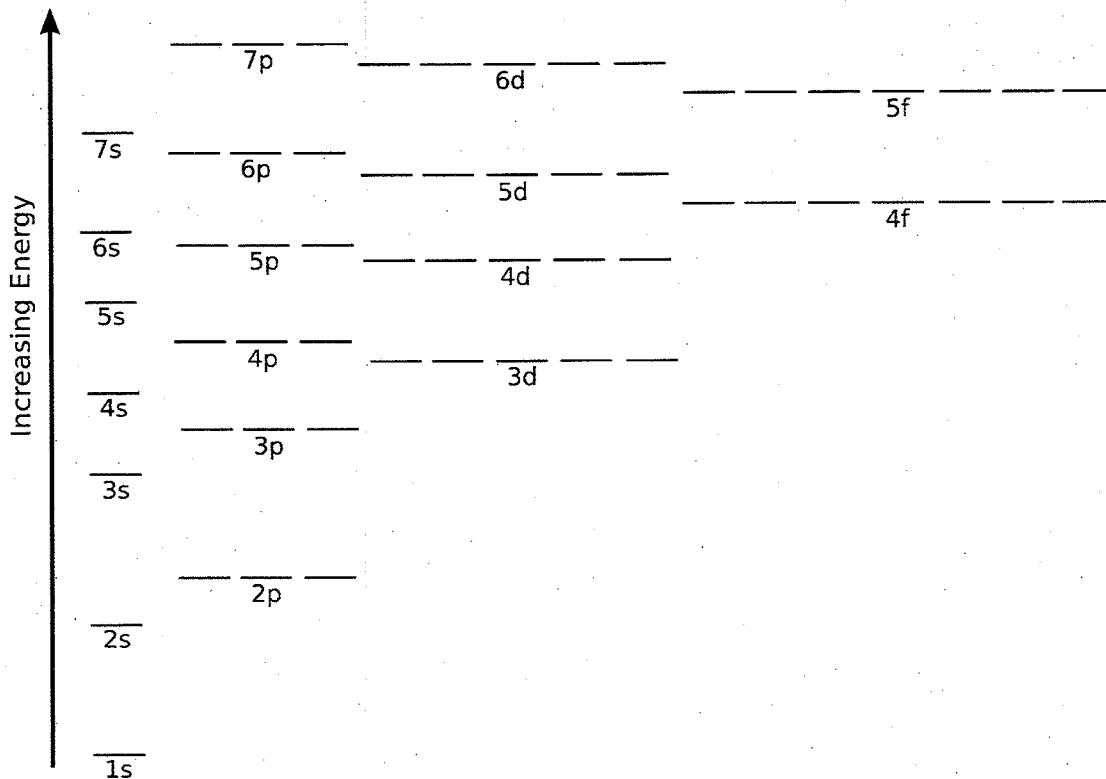
Quantum Mechanical Model Practice Problems

1. What are the possible values for the Principal Quantum Number?
2. What is the significance of the Principal Quantum Number?
3. Which energy level has the greatest energy; $n = 1$ or $n = 2$?
4. What is the significance of the Secondary Quantum Number?
5. How many subshells are in the
 - a) 1st energy level? b) 2nd energy level?
 - c) 3rd energy level c) 10th energy level?
6. Name the subshells in the following energy levels:
 - a) $n = 1$ b) $n = 2$ c) $n = 3$ d) $n = 4$
7. Within the 3rd energy level, which subshell has the
 - a) greatest energy b) lowest energy
8. What is the shape of the orbitals in the
 - a) s sublevel b) p sublevel
9. What is the significance of the Magnetic Quantum Number?
10. How many orbitals are in the following subshells?
 - a) s b) p c) d d) f
11. How many electrons fit into one orbital?
12. How many electrons fit into each of the following subshells?
 - a) s b) p c) d d) f
13. How many electrons can fit into the following energy levels?
 - a) $n = 1$ b) $n = 2$ c) $n = 3$ d) $n = 4$ e) $n = 5$
14. What is an orbital?

3.5 ATOMIC STRUCTURE AND THE PERIODIC TABLE

Energy Level Diagrams

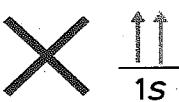
- chemists use the energy level diagram as well as electron configuration notation to represent which energy level, subshell, and orbital are occupied by electrons in any particular atom



Pauli Exclusion Principle

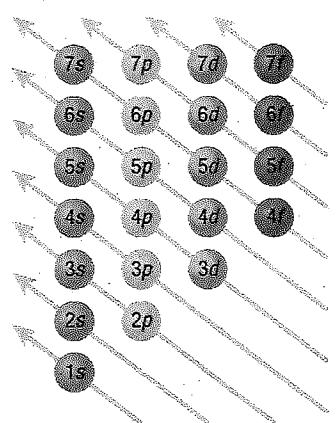
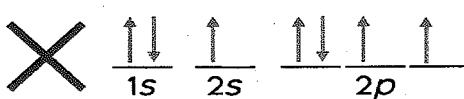
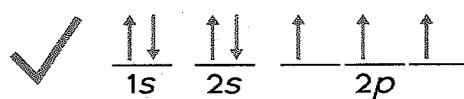
No two electrons in an atom can have the same four quantum numbers

- this means only electrons of _____ can occupy an orbital
- this can be shown using _____ of _____

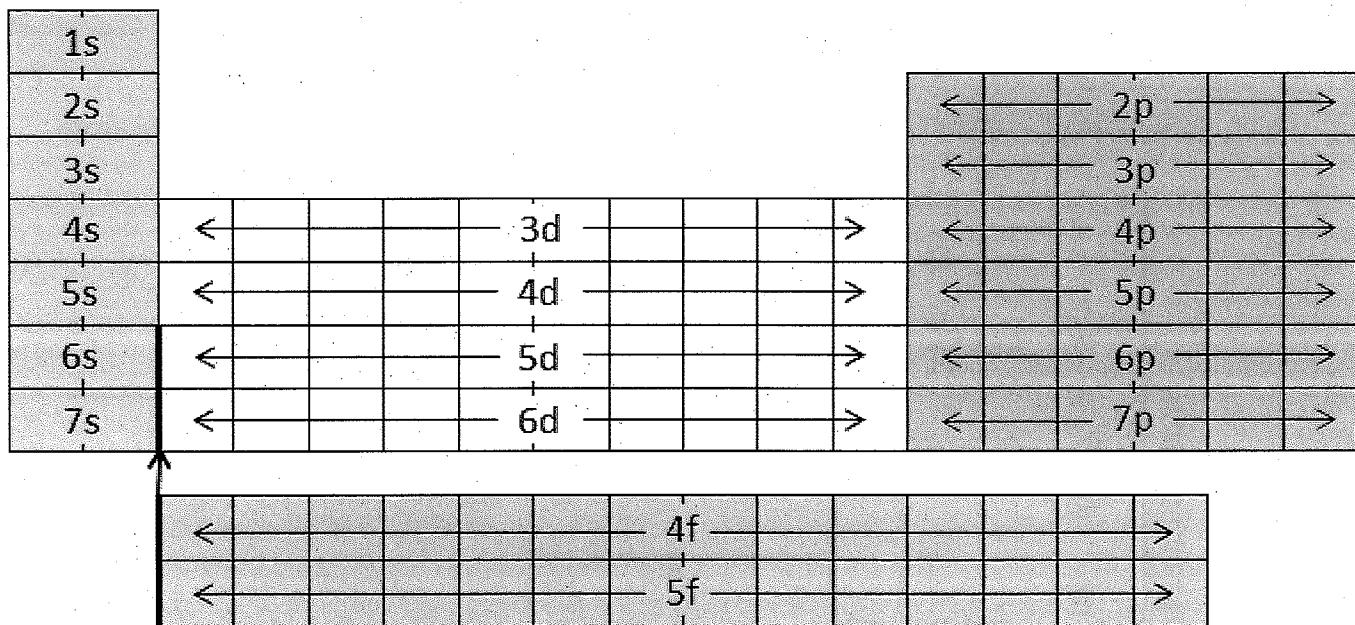


Aufbau Principle

- Place electrons into the orbitals in order of increasing energy level.
- Each subshell of orbitals in the same energy level must be completely filled before proceeding to the next orbital or series of orbitals.

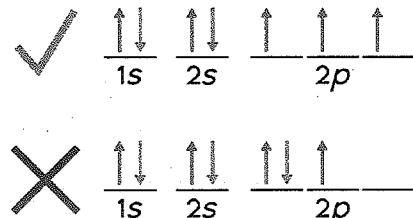


Periodic Table Organized by Orbitals

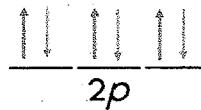


Hund's Rule

- Whenever electrons are added to orbitals of the same energy sublevel, each orbital receives one electron before pairing occurs.



- When electrons are added singly to separate orbitals of the same energy, the electrons must all have the same spin.



Examples: Draw the energy level diagram for

a) carbon

b) potassium

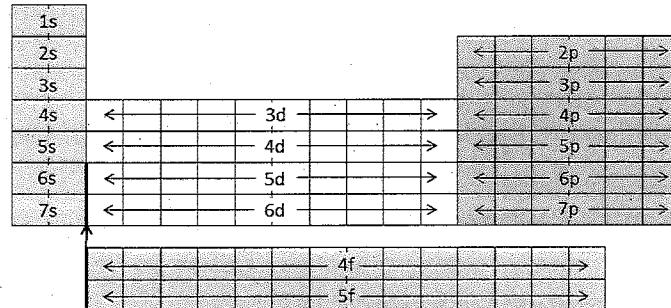
c) nickel

ELECTRON CONFIGURATION: An atom's electron configuration is a shorthand notation that shows the number and arrangement of electrons in orbitals

Example: What is the electron configuration for

a) chlorine atom

b) lead atom



Example: Write the electron configuration for the following:

(a) tin _____

(b) tin (II) ion _____

(c) tin (IV) ion _____

Example: Identify the element whose atoms have the electron configuration:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^4$ _____

Ground State vs. Excited State

- when a _____, it enters what is known as an _____
- the excited state can commonly take the form of the _____ from its original energy level or orbital, also known as the _____, to an empty orbital of a _____ shell that is further away from the nucleus.

Example: Nitrogen at its ground state: _____

Nitrogen at an excited state: _____

Isoelectronic Electron Configurations

- atoms and ions that possess the _____ of _____ would have the same electron configuration, and are said to be _____
- elements coming before and after noble gases form _____ ions by gaining or losing _____ until they have the same number of electrons as the _____
- N^{3-} , O^{2-} , F^- , Ne , Na^+ , Mg^{2+} , and Al^{3+} are isoelectronic as they all have identical electron configurations.

Li 3 6.94	Be 4 9.01	B 5 10.81	C 6 12.01	N 7 14.01	O 8 16.00	F 9 19.00	Ne 10 20.18
Na 11 22.99	Mg 12 24.31	Al 13 26.98	Si 14 28.09	P 15 30.97	S 16 32.06	Cl 17 35.45	Ar 18 39.95

Shorthand Electron Configurations: Write the chemical symbol for the previous noble gas in square brackets, then you continue with the configuration of the next energy level that is being filled

Example: a) Write the electron configuration for the **silver** atom

b) Write the shorthand form of the electron configuration for the **silver** atom

Exceptions in Period 4/11

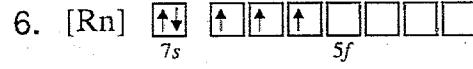
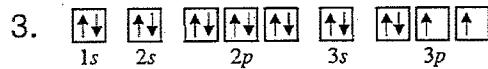
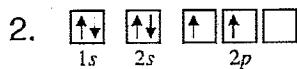
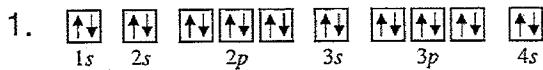
Cr

Cu

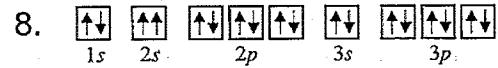
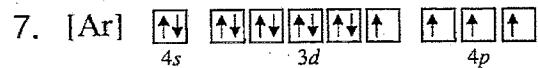
These electron configurations do not conform to the regular filling pattern but allow the atoms to achieve greater stability.

ENERGY LEVEL DIAGRAM PRACTICE

Write the name and symbol for the elements with the following orbital diagrams.



There is an error with each of the following orbital diagrams. Explain the error.



Write the orbital diagrams for the following:

Boron	Sodium	Phosphorus	Oxygen	Nitrogen

ELECTRON CONFIGURATION PRACTICE

Write the name and symbol for the atoms with the following electron configurations.

1. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
2. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
3. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^7$
4. $1s^2 2s^2 2p^6 3s^2 3p^1$
5. $[Rn]7s^2 5f^9$
6. $[Xe] 6s^2 4f^4 5d^{10} 6p^2$

Write complete electron configurations for the following substances.

7. nitrogen
8. magnesium
9. niobium
10. nickel
11. tin
12. chlorine

Write abbreviated electron configurations for the following elements.

13. arsenic
14. thulium
15. rubidium
16. einsteinium
17. platinum
18. molybdenum
19. sulfur
20. zirconium
21. argon
22. iron
23. polonium
24. bohrium

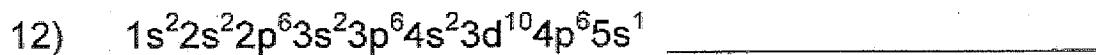
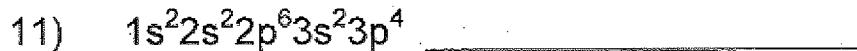
In the space below, write the unabbreviated electron configurations of the following elements:

- 1) sodium _____
- 2) iron _____
- 3) bromine _____
- 4) barium _____
- 5) neptunium _____

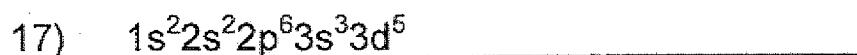
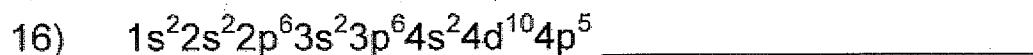
In the space below, write the abbreviated electron configurations of the following elements:

- 6) cobalt _____
- 7) silver _____
- 8) tellurium _____
- 9) radium _____
- 10) lawrencium _____

Determine what elements are denoted by the following electron configurations:



Determine which of the following electron configurations are not valid:



Atomic Theory Scientist Practice

Multiple Choice:

1. When assigning electrons in orbitals, each electron will first half-fill all the orbitals of the same energy before pairing with another electron in a half-filled orbital is known as:

- (A) Aufbau Principle (B) Dalton's Theory (C) Hund's Rule (D) Pauli Exclusion Principle

2. Which of the following states that electrons enter orbitals of lowest energy first?

- (A) Aufbau Principle (B) Dalton's Theory (C) Hund's Rule (D) Pauli Exclusion Principle

3. If two electrons occupy the same suborbital they must have opposite spins is known as:

- (A) Aufbau Principle (B) Dalton's Theory (C) Hund's Rule (D) Pauli Exclusion Principle

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SOLID SPHERE MODEL

PLUM PUDDING MODEL

NUCLEAR MODEL

PLANETARY MODEL

QUANTUM MODEL

JOHN DALTON



1803

Dalton drew upon the Ancient Greek idea of atoms (the word 'atom' comes from the Greek 'atomos' meaning indivisible). His theory stated that atoms are indivisible, those of a given element are identical, and compounds are combinations of different types of atoms.

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Schrödinger stated that electrons do not move in set paths around the nucleus, but in waves. It is impossible to know the exact location of the electrons; instead, we have 'clouds of probability' called orbitals, in which we are more likely to find an electron.

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RECOGNISED ELECTRONS AS COMPONENTS OF ATOMS



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ATOMS AREN'T INVISIBLE - THEY'RE COMPOSED FROM SUBATOMIC PARTICLES



NO NUCLEUS, DIDN'T EXPLAIN LATER EXPERIMENTAL OBSERVATIONS



DO NOT EXPLAIN WHY ELECTRONS REMAIN IN ORBIT AROUND THE NUCLEUS



Moving electrons should emit energy and collapse into the nucleus model did not work well for heavier atoms



STILL RICELY ACCEPTED AS THE MOST ACCURATE MODEL OF THE ATOM



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3.4 QUANTUM NUMBERS

Schrodinger's wave equation describes the **quantized energies of the electron** in an atom and the wave function that determines **the probability of finding electrons** in various regions of the atom.

Solving the wave equation gives a set of _____, and their corresponding _____. Electrons are found in these orbitals, which have various 3D shapes and sizes.

Quantum numbers are like addresses for locating the **position of an electron** by its city street, number and apartment number.

An orbital is described by a set of four quantum numbers:

- Principal quantum number, _____ : _____
- Angular momentum (secondary) quantum number, _____ : _____
- Magnetic quantum number, _____ : _____
- Spin quantum number, _____ : _____

Principal Quantum Number, n

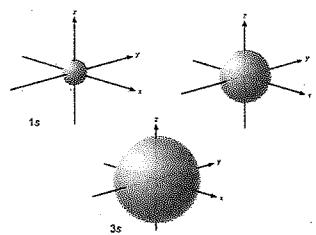
- indicates the relative _____ and _____ of atomic orbitals
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- $n = 1, 2, 3, 4 \dots$
- as n _____, orbital size _____
- maximum number of electrons per energy level: _____

Angular Momentum (Secondary) Quantum Number, ℓ

- energy sublevels are contained within the principal energy levels
- $\ell = \text{_____}$ each energy sublevel (s, p, d , or f) relates to orbitals of different shape

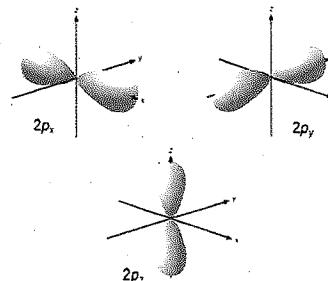
1. s orbitals

- $\ell = \text{_____}$
- each energy level has _____ s orbital
- s orbitals can hold _____



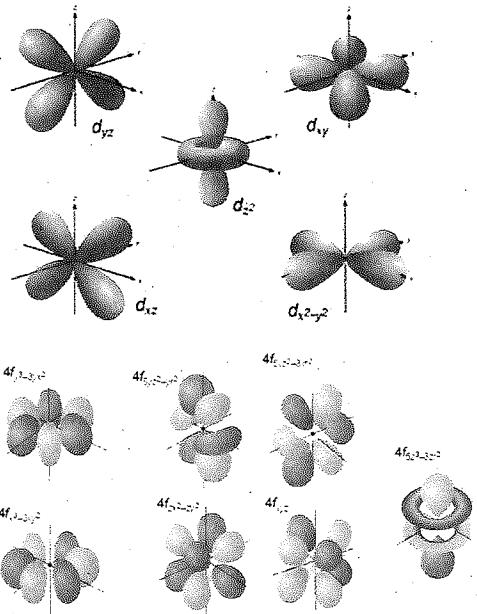
2. p orbitals

- $\ell = \text{_____}$
- p orbitals have greater energy than s orbitals
- _____ p orbitals per subshell
- each p orbital can hold _____ for a total _____ in the p subshell
- _____



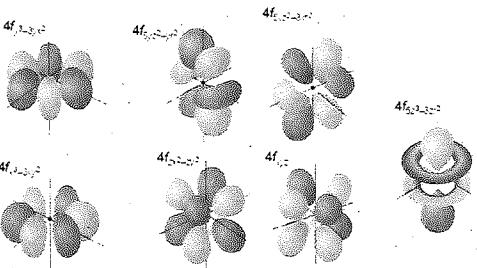
3. d orbitals

- $\ell = \underline{\hspace{2cm}}$
- d orbitals have greater energy than p orbitals
- $\underline{\hspace{2cm}}$ d orbitals in each subshell
- each d orbital can hold $\underline{\hspace{2cm}}$, for a total $\underline{\hspace{2cm}}$ in the d subshell



4. f Orbitals

- $\ell = \underline{\hspace{2cm}}$
- f orbitals have greater energy than d orbitals
- $\underline{\hspace{2cm}}$ f orbitals in each subshell
- each f orbital can hold $\underline{\hspace{2cm}}$, for a total $\underline{\hspace{2cm}}$ in the f subshell



Magnetic Quantum Number, m_l

- represents the orientation of the subshells
- has values of $\underline{\hspace{2cm}}$ to $\underline{\hspace{2cm}}$

$$s = \begin{array}{|c|} \hline \text{ } \\ \hline \end{array} 0$$

$$p = \begin{array}{|c|c|c|} \hline \text{ } & \text{ } & \text{ } \\ \hline -1 & 0 & 1 \\ \hline \end{array}$$

$$d = \begin{array}{|c|c|c|c|c|} \hline \text{ } & \text{ } & \text{ } & \text{ } & \text{ } \\ \hline -2 & -1 & 0 & 1 & 2 \\ \hline \end{array}$$

$$f = \begin{array}{|c|c|c|c|c|c|c|} \hline \text{ } & \text{ } \\ \hline -3 & -2 & -1 & 0 & 1 & 2 & 3 \\ \hline \end{array}$$

Spin Quantum Number, m_s

- represents the spin of the electron
- has values of $\underline{\hspace{2cm}}$ or $\underline{\hspace{2cm}}$

Example 1: For a given atom, $n = 3$:

(a) What are all possible values of l for this atom?

(b) What are all possible values for m_l for this atom?

(c) What is the maximum number of electrons this atom could have in this energy level?

Example 2: Which of the following sets of Quantum Numbers could NOT exist?

(a) $n = 3$ $l = 2$ $m_l = -1$ $m_s = +\frac{1}{2}$

(b) $n = 2$ $l = 2$ $m_l = 0$ $m_s = -\frac{1}{2}$

(c) $n = 3$ $l = 1$ $m_l = -2$ $m_s = +\frac{1}{2}$

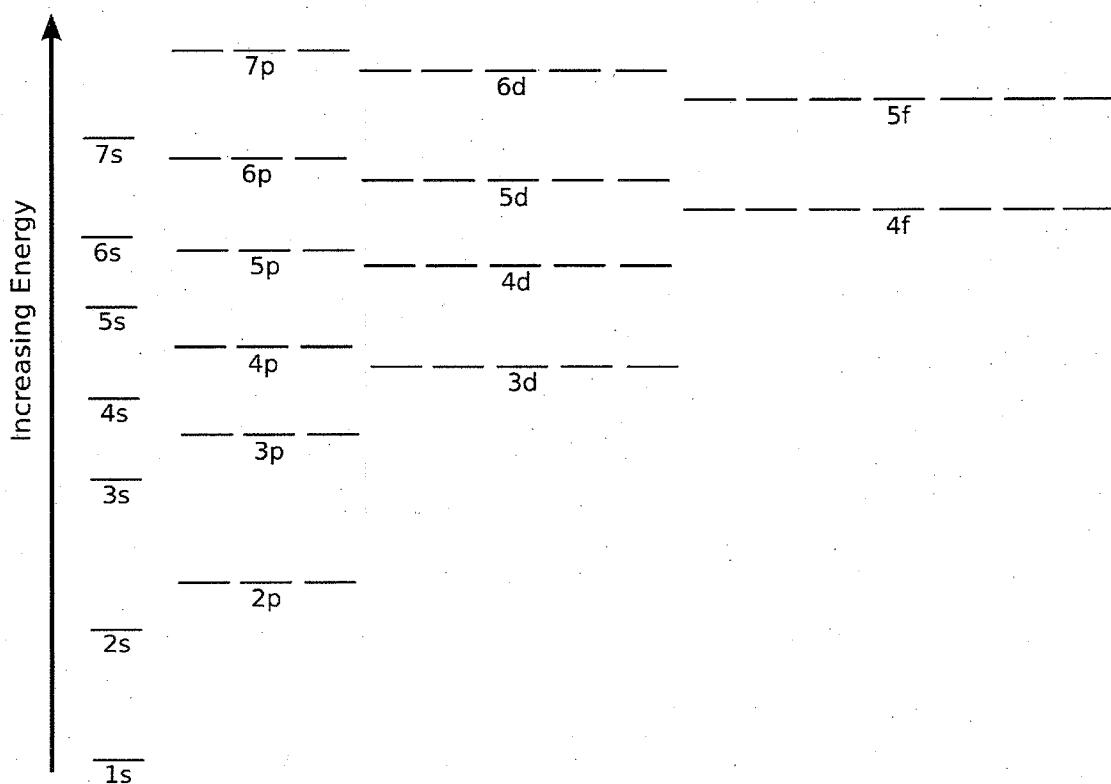
Quantum Mechanical Model Practice Problems

1. What are the possible values for the Principal Quantum Number?
2. What is the significance of the Principal Quantum Number?
3. Which energy level has the greatest energy; $n = 1$ or $n = 2$?
4. What is the significance of the Secondary Quantum Number?
5. How many subshells are in the
 - a) 1st energy level? b) 2nd energy level?
 - c) 3rd energy level d) 10th energy level?
6. Name the subshells in the following energy levels:
 - a) $n = 1$ b) $n = 2$ c) $n = 3$ d) $n = 4$
7. Within the 3rd energy level, which subshell has the
 - a) greatest energy b) lowest energy
8. What is the shape of the orbitals in the
 - a) s sublevel b) p sublevel
9. What is the significance of the Magnetic Quantum Number?
10. How many orbitals are in the following subshells?
 - a) s b) p c) d d) f
11. How many electrons fit into one orbital?
12. How many electrons fit into each of the following subshells?
 - a) s b) p c) d d) f
13. How many electrons can fit into the following energy levels?
 - a) $n = 1$ b) $n = 2$ c) $n = 3$ d) $n = 4$ e) $n = 5$
14. What is an orbital?

3.5 ATOMIC STRUCTURE AND THE PERIODIC TABLE

Energy Level Diagrams

- chemists use the energy level diagram as well as electron configuration notation to represent which energy level, subshell, and orbital are occupied by electrons in any particular atom



Pauli Exclusion Principle

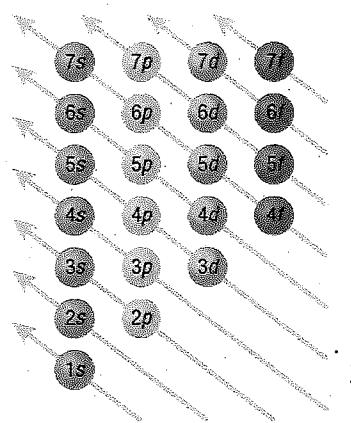
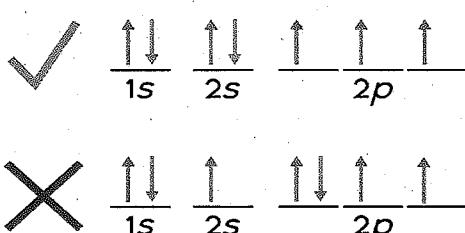
No two electrons in an atom can have the same four quantum numbers

- this means only electrons of _____ can occupy an orbital
- this can be shown using _____ of _____

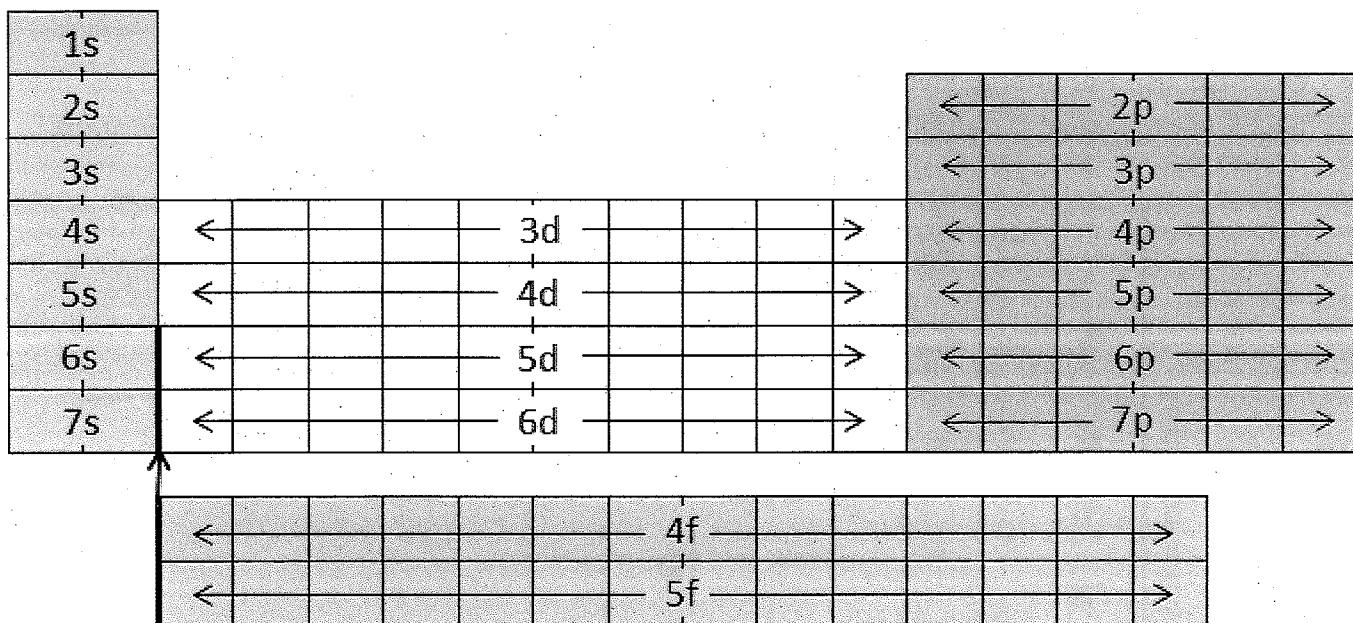


Aufbau Principle

- Place electrons into the orbitals in order of increasing energy level.
- Each subshell of orbitals in the same energy level must be completely filled before proceeding to the next orbital or series of orbitals.

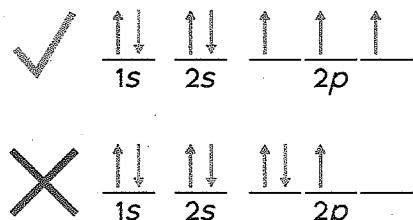


Periodic Table Organized by Orbitals

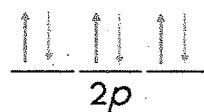


Hund's Rule

- Whenever electrons are added to orbitals of the same energy sublevel, each orbital receives one electron before pairing occurs.



2. When electrons are added singly to separate orbitals of the same energy, the electrons must all have the same spin.



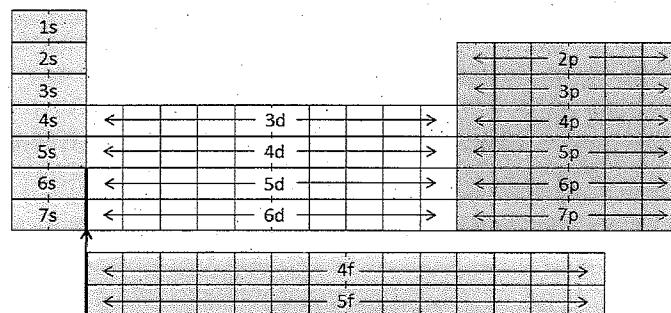
Examples: Draw the energy level diagram for

ELECTRON CONFIGURATION: An atom's electron configuration is a shorthand notation that shows the number and arrangement of electrons in orbitals

Example: What is the electron configuration for

a) chlorine atom

b) lead atom



Example: Write the electron configuration for the following:

(a) tin _____

(b) tin (II) ion _____

(c) tin (IV) ion _____

Example: Identify the element whose atoms have the electron configuration:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^4$ _____

Ground State vs. Excited State

- when a _____, it enters what is known as an _____
- the excited state can commonly take the form of the _____ from its original energy level or orbital, also known as the _____, to an empty orbital of a _____ shell that is further away from the nucleus.

Example: Nitrogen at its ground state: _____

Nitrogen at an excited state: _____

Isoelectronic Electron Configurations

- atoms and ions that possess the _____ of _____ would have the same electron configuration, and are said to be _____
- elements coming before and after noble gases form _____ ions by gaining or losing _____ until they have the same number of electrons as the _____
- N^{3-} , O^{2-} , F^- , Ne , Na^+ , Mg^{2+} , and Al^{3+} are isoelectronic as they all have identical electron configurations.

Li 6.94	Be 9.01	B 10.81	C 12.01	N 14.01	O 16.00	F 19.00	Ne 20.18
Na 22.99	Mg 24.31	Al 26.98	Si 28.09	P 30.97	S 32.06	Cl 35.45	Ar 39.95

Shorthand Electron Configurations: Write the chemical symbol for the previous noble gas in square brackets, then you continue with the configuration of the next energy level that is being filled

Example: a) Write the electron configuration for the **silver** atom

b) Write the shorthand form of the electron configuration for the **silver** atom

Exceptions in Period 4/11

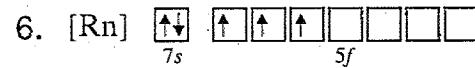
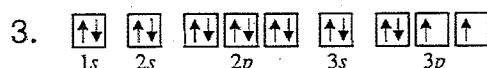
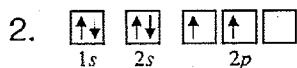
Cr

Cu

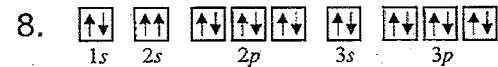
These electron configurations do not conform to the regular filling pattern but allow the atoms to achieve greater stability.

ENERGY LEVEL DIAGRAM PRACTICE

Write the name and symbol for the elements with the following orbital diagrams.



There is an error with each of the following orbital diagrams. Explain the error.



Write the orbital diagrams for the following:

Boron	Sodium	Phosphorus	Oxygen	Nitrogen
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ELECTRON CONFIGURATION PRACTICE

Write the name and symbol for the atoms with the following electron configurations.

1. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
4. $1s^2 2s^2 2p^6 3s^2 3p^1$
2. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
5. $[Rn]7s^2 5f^9$
3. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^7$
6. $[Xe] 6s^2 4f^{14} 5d^{10} 6p^2$

Write complete electron configurations for the following substances.

7. nitrogen
10. nickel
8. magnesium
11. tin
9. niobium
12. chlorine

Write abbreviated electron configurations for the following elements.

13. arsenic
19. sulfur
14. thulium
20. zirconium
15. rubidium
21. argon
16. einsteinium
22. iron
17. platinum
23. polonium
18. molybdenum
24. bohrium

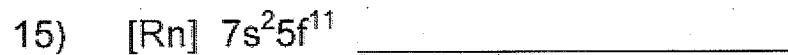
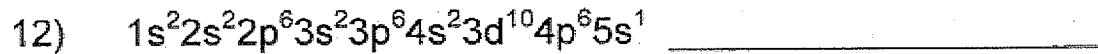
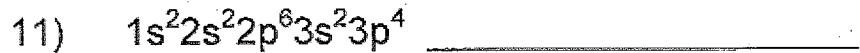
In the space below, write the unabbreviated electron configurations of the following elements:

- 1) sodium _____
- 2) iron _____
- 3) bromine _____
- 4) barium _____
- 5) neptunium _____

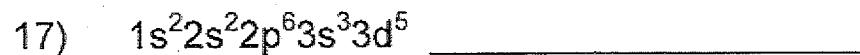
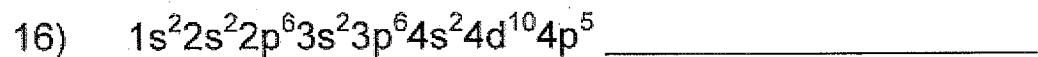
In the space below, write the abbreviated electron configurations of the following elements:

- 6) cobalt _____
- 7) silver _____
- 8) tellurium _____
- 9) radium _____
- 10) lawrencium _____

Determine what elements are denoted by the following electron configurations:



Determine which of the following electron configurations are not valid:



Atomic Theory Scientist Practice

Multiple Choice:

1. When assigning electrons in orbitals, each electron will first half-fill all the orbitals of the same energy before pairing with another electron in a half-filled orbital is known as:

- (A) Aufbau Principle (B) Dalton's Theory (C) Hund's Rule (D) Pauli Exclusion Principle

2. Which of the following states that electrons enter orbitals of lowest energy first?

- (A) Aufbau Principle (B) Dalton's Theory (C) Hund's Rule (D) Pauli Exclusion Principle

3. If two electrons occupy the same suborbital they must have opposite spins is known as:

- (A) Aufbau Principle (B) Dalton's Theory (C) Hund's Rule (D) Pauli Exclusion Principle