CP Chemistry Final Exam Review Topics & Problems

Chapter 2- Analyzing Data

2.1- Units & Measurement- Converting units in the metric system.

2.2- Scientific Notation & Dimensional Analysis- Converting units between the metric system and the English system.

Chapter 3-Matter- Properties & Changes

3.1- Properties of Matter- Descriptions of phases of matter and chemical & physical properties.

Example-

Calculate the density of a solid which has a mass of 18.325 grams. The volume was determined by putting the solid in a graduated cylinder. Before the solid was put in the cylinder, the volume of water was 10.00 mL. After the solid was put in the cylinder, the volume of water and cylinder was 18.60 mL.

$D = \frac{M}{V} = \frac{18.325\text{g}}{(18.6-10)} = \boxed{2.13\ \text{g/mL}}$

3.2- Changes in Matter- Differences between chemical & physical changes.

3.3- Mixtures in Matter- Heterogeneous & Homogeneous mixtures.

3.4- Elements & Compounds- Elements are simplest materials & compounds are chemical combinations of 2 or more elements in definite proportions.

Chapter 4- The Structure of the **Atom**

4.1- Early Ideas about Matter- Historical description of the thoughts of what made up the atom.

4.2- Defining the Atom- Identifying the function & location of protons, neutrons & electrons.

4.3- How Atoms Differ- Identifying the atomic number (# of protons = # of electrons), atomic mass (# of protons + # of neutrons), calculating the average atomic mass of an element.

Example- Cerium has 2 common isotopes, Cerium-140 (Abundance: 94.5%) and Cerium-142 (Abundance: 5.5%). Calculate the average atomic mass of cerium. How many protons, neutrons & electrons are in both isotopes of cerium?

Avg At Mass = $(140)(.945) + (142)(.055) = 132.3 + 7.81 = 140.11$ g/mol

Cerium-140 - At #58 - 58 protons, 58 electrons, 82 neutrons

Cerium-142 - At #58, 58 protons, 58 electrons, 84 neutrons

Chapter 5- Electrons in Atoms

5.3- Electron Configuration- writing ground state electron configuration of atoms and ions. (1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, etc.)

Example- Write the complete electron configuration and noble gas configuration for Chromium, Sulfur & Molybdenum.

Cr - At #24 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ $[Ar]4s^2 3d^4$

S - At #16 $1s^2 2s^2 2p^6 3s^2 3p^4$ $[Ne]3s^2 3p^4$

Mo - At #42 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^4$ $[Kr]5s^2 4d^4$

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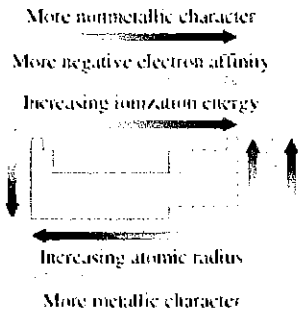
Chapter 6- The Periodic Table & Periodic Law

6.2- Classification of elements- periods (rows) and families (columns), valence electrons = # of electrons in outer shell = the column in the s and p blocks of table.

6.3- Periodic Trends-

Atomic Radius- increase down families, decrease across period
Ionic Radius- increase down families, decrease across period
Ionization Energy- decrease down families, increase across period
Electronegativity- decrease down families, increase across period



Chapter 7- Ionic Compounds & Metals

7.1- Ion Formation- An ion is an atom that either gains or loses electrons. A cation is positively charged. An anion is negatively charged. They either gain or lose electrons to have a full outer shell of electrons (s and p shells).

7.3- Names and Formulas for Ionic Compounds- Ionic compounds are combinations of a metal cation and a nonmetal anion and have no prefixes in their names. You figure out the formula from the name by balancing out the charges of the cations and anions so the total charge of the compound is zero. Use the criss cross rule.

Example- Write the formulas for the following ionic compounds-

a. Potassium Carbonate K^{+1} CO_3^{-2} K_2CO_3
b. Sodium Hydroxide Na^{+1} OH^{-} $NaOH$
c. Sodium Chloride Na^{+1} Cl^{-} $NaCl$
d. Potassium Bromide K^{+1} Br^{-1} KBr
e. Copper (II) Hydroxide Cu^{+2} OH^{-1} $Cu(OH)_2$
f. Iron (III) Sulfate Fe^{+3} SO_4^{-2} $Fe_2(SO_4)_3$
g. Iron (II) Sulfate Fe^{+2} SO_4^{-2} $FeSO_4$
h. Potassium Sulfate K^{+1} SO_4^{-2} K_2SO_4
i. Ammonium Chloride NH_4^{+1} Cl^{-1} NH_4Cl
j. Ammonium Nitrate NH_4^{+1} NO_3^{-1} NH_4NO_3
k. Silver Nitrate Ag^{+1} NO_3^{-1} $AgNO_3$
l. Magnesium Chloride Mg^{+2} Cl^{-1} $MgCl_2$

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Chapter 9- Chemical Reactions

9.1-Reactions & Equations- balancing equations (total molecules of reactants = total molecules of products) using coefficients.

9.2-Classifying Chemical Reactions- The basic types of reactions- synthesis, decomposition, single replacement, double replacement, combustion.

Example- Balance the following reactions and identify the reaction type.

a. 1 K_2CO_3 + 2 C → 3 CO + 2 K

b. 2 H_2 + 1 O_2 → 2 H_2O - Combination/Synthesis

c. 1 $Fe_2(SO_4)_3$ + 6 KOH → 3 K_2SO_4 + 2 $Fe(OH)_3$ - Double Replacement

d. 4 Sb + 3 O_2 → 1 Sb_4O_6 - Combination/Synthesis/Combustion

e. 2 K + 1 Br_2 → 2 KBr - Combination/Synthesis

f. 1 HNO_3 + 1 KOH → 1 H_2O + 1 KNO_3 - Double Replacement / Acid-Base

g. 2 HCl + 1 Zn → 1 H_2 + $ZnCl_2$ Single Replacement

Chapter 10-The Mole

Mole finding tool- grams ⟷ (×MM / ÷MM) moles ⟷ (÷6.02×10^{23} / ×6.02×10^{23}) particles

Example- How many moles are in 210 g of water? H_2O MM = 18

$210\ g\ H_2O \div 18\ g/mole = 11.67\ mole\ H_2O$

How many moles are in 146 g of CH_4? MM = 16

$146\ g \div 16\ g/mole = 9.125\ mole\ CH_4$

How many atoms are in 80 grams of calcium? MM = 40

$80\ g\ Ca \div 40\ g/mole = 2\ mole \times 6.02\times10^{23} = 1.204\times10^{24}\ atoms$

How many grams of benzene (C_6H_6) are needed to make 0.75 mol of benzene? MM = 78

$.75\ mol\ C_6H_6 \times 78\ g/mole = 58.5\ g$

Chapter 11- Stoichiometry

Mathematical process of using mole ratios to predict the amount of reactant needed in a reaction or the amount of product in a reaction. You cannot compare grams of substances, it must be moles.

Steps- 1. Balance equation
2. Convert grams of given to moles of given (if necessary).
3. Convert moles of given to moles of needed (use mole ratio from balanced equation).
4. Convert moles of needed to grams of needed (if necessary).

Example- How many grams of water are produced when 2.5 mol oxygen reacts with enough hydrogen? $2H_2 + O_2 \rightarrow 2H_2O$

$$2.5\ mol\ O_2 \left(\frac{2\ mol\ H_2O}{1\ mol\ O_2}\right)\left(\frac{18\ g}{1\ mol\ H_2O}\right) = 90\ g\ H_2O$$

How many moles of Cl_2 react with 6.0 mol of P_4? Balanced equation: $P_4 + 10Cl_2 \rightarrow 4PCl_5$

$$6\ mol\ P_4 \left(\frac{10\ mol\ Cl_2}{1\ mol\ P_4}\right) = 60\ mol\ Cl_2$$

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How many grams of silver nitrate are needed to react with 3.93 grams of magnesium chloride?

$AgNO_3$ MM = 170 g/mol $MgCl_2$ MM = 95.3

$$2AgNO_3 + MgCl_2 \longrightarrow 2AgCl + Mg(NO_3)_2$$

$$3.93\ g\ MgCl_2 \left(\frac{1\ mol\ MgCl_2}{95.3\ g}\right)\left(\frac{2\ mol\ AgNO_3}{1\ mol\ MgCl_2}\right)\left(\frac{170\ g}{1\ mol\ AgNO_3}\right) = 14.0\ g\ AgNO_3$$

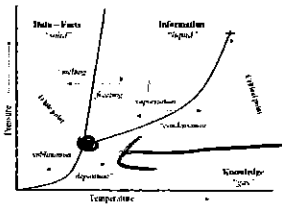
Chapter 12 & 13- Gases

12.1- Effusion- The heavier the gas, the slower it will travel as its particles spread through the air (example- perfume or skunk odor)

Example- What is the faster gas, krypton or neon, at the same temperature and pressure?

MM = 83.8 MM = 20.2 Neon is faster, it has a lower MM.

12.4- Phase changes and Phase Diagrams-



P

T

Temp + Pressure

Triple Point - where the S, L + Gas Phase all exist in equilibrium

Chapter 13- Gases & Gas Laws

13.1 Basic Gas Laws

Boyle's Law $P_1V_1 = P_2V_2$

Gay-Lussac's Law $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Charles' Law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Combined Gas Law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Example: Use the appropriate gas law to calculate the missing value in each of the following:

a. $V_1 = 2.0$ L $P_1 = 0.82$ atm $V_2 = 1.0$ L $P_2 =$

$P_1V_1 = P_2V_2$ $(.82)(2) = P_2(1)$ $P_2 = 1.64$ atm

b. $V_1 = 750$ mL $T_1 = 298$ K $V_2 =$ $T_2 = 323$ K

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $\frac{750\ mL}{298\ K} = \frac{V_2}{323\ K}$ $V_2 = 812.9$ mL

c. $P_1 = 250$ mm Hg $T_1 =$ $P_2 = 400$ mm Hg $T_2 = 298$ K

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ $\frac{250\ mm}{T_1} = \frac{400\ mm}{298\ K}$ $T_1 = 186.25$ K

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13.2- Ideal Gas Law

PV = nRT Molar Volume of a gas at STP = 22.4 L/mole

R = 0.0821 Latm/molK STP = standard temp (273K) and pressure (1atm)

Example-

Calculate the volume of a gas when a 0.650 mol sample at 15.0 °C and has a pressure of 1.1 atm exerted on it.

$PV = nRT$ $V = \frac{nRT}{P} = \frac{(.650\ mol)(.0821\ \frac{L\cdot atm}{mol\cdot K})(288\ K)}{1.1\ atm} = 13.97\ L$

A 500.0 mL container is filled with 2.20 g of carbon dioxide gas, CO_2, at 400 K. What pressure is exerted by the gas? .5 L

$n = moles = \frac{grams}{MM} = \frac{2.2g}{44g/mol} = .05\ mol$ MM = 44 g/mol

$PV = nRT$ $P = \frac{nRT}{V} = \frac{(.05\ mol)(.0821\ \frac{L\cdot atm}{mol\cdot K})(400K)}{.5\ L} = 3.284\ atm$

What is the volume of 2.50 mol CO_2 at STP?

$2.50\ mol\ CO_2 \left(\frac{22.4\ L}{mole}\right) = 56\ L$ or $V = \frac{nRT}{P} = \frac{(2.5)(.0821)(273)}{(1)} = 56\ L$

How many moles of CO_2 are present in 11.2 L at STP?

$11.2\ L\ CO_2 \left(\frac{1\ mol}{22.4\ L}\right) = .5\ mol\ CO_2$ or $n = \frac{PV}{RT} = \frac{(1)(11.2L)}{(.0821)(273)} = .5\ mol$

Chapter 14- Mixtures & Solutions

14.1- Definitions

Solution - ~~Heterogeneous~~ Homogeneous mixture

Solute - stuff being dissolved

Solvent - stuff doing the dissolving

Suspension - Heterogeneous mixture that settles when left undisturbed

Colloid - Heterogeneous mixture that scatters light

Supersaturated Solution - A solution that contains MORE than the max amount of solute

Saturated Solution - A solution that contains the max amount of solute.

Unsaturated Solution - A solution that contains LESS than the max amount of solute

14.2- Measurement of Concentration

% Mass

$\%\ Mass = \frac{mass\ part}{mass\ whole} \times 100$

Molality (*m*)

$m = \frac{moles\ of\ solute}{kg\ solvent}$

Molarity (M)

$M = \frac{moles\ solute}{L\ solution}$

Dilution $M_1V_1 = M_2V_2$

moles 1 = moles 2

14.4- Colligative Properties- properties that depend on the amount of solute as opposed to the type of solute- boiling point elevation ($\Delta t_b = k_b m$) and freezing point depression ($\Delta t_f = k_f m$)

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Examples-

What is the molarity of a solution that contains 5.8 grams of NaCl dissolved in 20 mL of water?

$5.8\text{ g NaCl} \div 58.5\text{ g/mole} = .0991\text{ mol NaCl}$

$20\text{ mL} \div 1000 = .02\text{ L}$

$M = \frac{\text{moles}}{L} = \frac{.0991}{.02} = 4.95\text{ M}$

How many grams of HNO_3 is needed to dissolve in 50 mL of water to make a 0.25 M solution? MM=63

$\text{moles} = M \times L = (.25\text{ M})(.05\text{ L}) = .0125\text{ mol } HNO_3 \left(\frac{63\text{ g}}{1\text{ mole}}\right) = .7875\text{ g } HNO_3$

In the lab, you dissolve of 179 g of CH_4 in 1000 g of water. Determine the freezing and boiling points of the solution. $K_f = 1.86\ ^\circ C/m$ and $K_b = 0.52\ ^\circ C/m$

$179\text{ g } CH_4 \div 16\text{ g/mole} = 11.2\text{ mol}$

$m = \frac{\text{moles } CH_4}{\text{kg } H_2O} = \frac{11.2\text{ mol}}{1\text{ kg}} = 11.2\text{ m}$

$\Delta t_f = m \cdot k_f = (11.2\text{ m})(1.86\ ^\circ C/m) = 20.8^\circ C$

New FP = $-20.8^\circ C$

$\Delta t_b = m \times k_b = (11.2\text{ m})(.52\ ^\circ C/m) = 5.82^\circ C$

New BP = $105.8^\circ C$

Chapter 15- Energy

Measurement of energy- Joule or calorie (1 calorie = 4.184 J).

Calculating heat gained or lost in a substance: $Q = mc\Delta t$

Endothermic reaction = absorbs heat, Q = + number

Exothermic reaction = releases heat, Q = - number

In a reaction, $Q_{water} = -Q_{metal}$

Example- The temperature of 97.7 grams of copper is raised from 25 to 50 °C. The specific heat of copper is 0.38 J/g°C. How much heat does the copper absorb?

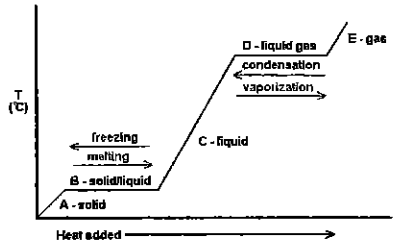
$Q = mc\Delta t = (97.7\text{ g})(.385\text{ J/g}\cdot^\circ C)(50-25) = 928\text{ J}$

110 grams of water is heated from 25 to 95 degree Celsius. The specific heat of water is 4.18 J/g°C. Is the reaction endo- or exothermic? How much heat is transferred?

$Q = mc\Delta t = (110\text{ g})(4.184\text{ J/g}\cdot^\circ C)(95-25) = 32,186\text{ J}$

absorbed - endotherm.

Phase Change Diagram



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Chapter 16- Kinetics

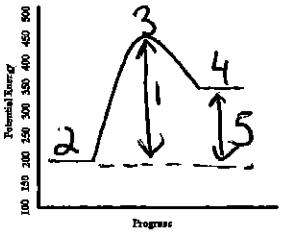
Collision Theory – particles react when they collide w/ enough energy + correct position

Activation Energy- minimum energy reactants need to have to become product

Catalyst- a chemical that speeds up a reaction by lowering the activation energy barrier. Does not get involved in the reaction

Factors that affect rates:

Example - particle size - smaller = faster temp - higher = faster
concentration - higher = faster Catalyst - use = faster



Match the appropriate number with the quantity it represents in the energy diagram to the left.

a. Reactants 2
b. Activated complex 3
c. Products 4
d. Activation energy 1
e. heat of reaction 5

Chapter 17-Equilibrium

Writing K expressions from a reaction and calculating the value for K. K = [products]/[reactants]

K > 1 favors products. K < 1 favors reactant.

LeChatelier's principle

Add reactant →
Add product ←
Remove reactant ←
Remove product →

Add heat - shift away from heat
Remove heat shift towards heat
Increase pressure shift toward side w/ less moles of gas
Decrease pressure shift toward side w/ more moles of gas

Example-

Given the equilibrium system with the equation $CO(g) + 2H_2(g) \leftrightarrows CH_3OH(g) + heat$

a. Write the K expression

$$K = \frac{[CH_3OH]}{[CO][H_2]^2}$$

b. If the equilibrium concentrations are [CO] = 2.0 M, $[H_2]$ = 4.0 M, $[CH_3OH]$ = 6.0 M, calculate the value for K.

$$K = \frac{[CH_3OH]}{[CO][H_2]^2} = \frac{(6)}{(2)(4)^2} = .1875$$

favors reactants (<1)

How would each of the following changes affect the equilibrium position of the system?

c. Adding CO to the system →
d. Cooling the system →
e. Adding a catalyst to the system No change
f. Removing CH_3OH from the system →
g. Increase the pressure of the system ←

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Chapter 18= Acids & Bases

When acids react with metals, Hydrogen gas is formed.
When acids react with bases, water and salt is formed.

Acids are hydrogen donors. Bases are hydrogen acceptors.

pH scale: pH < 7 = Acid; pH > 7 = Base; pH = 7 = Neutral

$[H^+] = 10^{-pH}$; $[OH^-] = 10^{-pOH}$; pH + pOH = 14

Indicators: phenolphthalein = pink in base; clear in acid/neutral.

Titration: A reaction that neutralizes and acid with a base or vice versa. Moles acid = moles base
$M_AV_A = M_BV_B$

Example:
What is $[OH^-]$ in an aqueous solution in which $[H^+] = 1 \times 10^{-3}$ M? What is pH? What is pOH?

pH = 3
pOH = 11 $[OH^-] = 1 \times 10^{-11}$

What is the molarity of nitric acid solution (HNO_3) if 43.33 mL of 0.100 M KOH solution is needed to neutralize 20.00 mL of the acid solution?

$M_AV_A = M_BV_B$
$(M_A)(43.33\text{mL}) = (.1\text{M})(20\text{mL})$ $M_A = .0462\text{M}$

Chapter 19= Oxidation/Reduction Reactions
How to determine oxidation #'s: Hydrogen is always +1, oxygen is usually -2, ions of elements in groups 1 and 2 are equal to their charge, neutral elements and diatomic molecules are 0.

Oxidation is Loss of electrons. Reduction is Gain of electrons. OIL RIG
A Redox reaction involves a transfer of electrons.

Example: Identify the species oxidized and the species reduced in the following redox reactions.