The following is a list of concepts and terms with which you should be well-acquainted. **THIS IS NOT A SUBSTITUTE FOR YOUR NOTES!!** This list of review topics is not meant to be exhaustive, it is only meant to help you identify the major areas we have discussed. **You are responsible for all material covered in lecture and in the text.**

**Chapter 4: Chemical Reactions**

The Nature of Solutions:
- speciation and stoichiometry associated with the dissolution of ionic vs. molecular compounds in water
- strong vs. weak vs. nonelectrolytes: what were the distinctions? what types of compounds fall in each category?
- solution composition: molarity (M) or molar concentration = mol solute/L sol’n
- calculations of molarity of solutions, molarity of ions in a solution, and using molarity as a conversion factor relating mol of solute and volume of solution
- calculations that involve molar concentration of solutions: dilution calculations, $M_i V_i = M_f V_f$ or you can think of it in terms of $M_{\text{conc}} V_{\text{conc}} = M_{\text{doll}} V_{\text{doll}}$
- stoichiometry of solutions in chemical reactions, volume of solution required for reaction with a given amount of reactant, or to form a given amount of product
- titrations, acid/base or redox – usually to determine something about the composition of the analyte

**DO NOT USE $M_i V_i = M_f V_f$ FOR REACTION STOICHIOMETRY OR TITRATION CALCULATIONS!!** You must demonstrate to me that you understand that a chemical reaction is occurring, and that you understand the mole ratios of reactants and products in that reaction.

Types of Chemical Reactions: We have discussed precipitation reactions, acid-base neutralization reactions, and oxidation-reduction reactions as well as calculations associated with each of them.
- Precipitation Reactions:
  - predict reaction products (and whether or not a reaction actually occurs) for a given set of reactants
  - apply solubility rules to predict whether products are soluble or insoluble
  - write the balanced chemical equation, the full ionic equation and net ionic equation for a precipitation reactions recognize the spectator ions.
- Acid/Base Neutralization Reactions:
  - Bronsted/Lowry definitions of acids and bases – proton donors or acceptors in aqueous solution
  - distinctions between strong and weak acids and bases in terms of definition and resulting speciation in aqueous solution
  - know the strong acids (HCl, HBr, HI, HNO$_3$, HClO$_4$, H$_2$SO$_4$) and strong bases (hydroxides of Grp 1A metals and Mg$^{2+}$, Ca$^{2+}$, Sr$^{2+}$, Ba$^{2+}$)
  - predict the products of a neutralization reaction; write the balanced chemical equation, the full ionic equation and the net ionic equation; identify spectator ions
  - acid/base properties of metal and non-metal oxides, carbonates, hydrogen carbonates, sulfates, hydrogen sulfates, sulfites, and hydrogen sulfides
- Redox reactions:
  - definitions of oxidation, reduction, oxidizing agent, and reducing agent
  - identify an oxidation/reduction reaction based on changes in oxidation states
  - write oxidation and reduction half reactions
  - determine oxidation numbers of elements within species and a reaction
  - identify oxidizing and reducing agents
  - balance redox reactions by the method of balancing half reactions for reactions in acidic or basic solution
  - recognize combination, decomposition, combustion and displacement reactions; predict products based on reaction type
- Activity Series used to predict hydrogen and metal displacement reaction products

**Chapter 5: The Gaseous State**
- fundamentals of gas behavior: how are changes in T, P, V or amount of substance inter-related? what is meant by the pressure exerted by a gas? what is the behavior of the individual particles in a gas sample?
- units of pressure and the conversions between them: mm Hg, Torr, atm, pascal (Pa) – the SI unit of pressure, kPa: 1 Torr = 1 mm Hg; 760 Torr = 1 atm; 101.3 kPa = 1 atm
- Gas Laws: equations that define how $P$, $T$, $V$, $n$ and even density and molar mass relate to one another; can be interpreted both qualitatively and quantitatively
  - Boyle's Law, $P_1 V_1 = P_2 V_2$ at constant $T$
  - Charles's Law, $V_1/T_1 = V_2/T_2$ at constant $P$
Combined Gas Law, \( P_1 V_1 / T_1 = P_2 V_2 / T_2 \) for a given amount of sample

Avogadro's Law, \( V \) is directly proportional to \( n \); molar volume of a gas \( = V/n \);

Molar Volume \( = 22.41 \text{ L/mol at STP} \)

STP (Standard Temperature and Pressure) conditions are defined as \( T = 0^\circ \text{C} \) and \( P = 1 \text{ atm} \)

Ideal Gas Law, \( PV = nRT \)

the relationships between density \( (d) \) and molar mass \( (M_n) \) of a gas derived from the ideal gas law: \( PM_n = dRT \)

\[
R = 0.08206 \text{ L•atm•K}^{-1}•\text{mol}^{-1}
\]

- Stoichiometry of reacting gases. Use stoichiometric relationships in the chemical equation to determine amounts or volumes of gases reacting at STP, and at other \( T \) and \( P \).

- Mixtures of non-reactive gases
  partial pressures, and Dalton's Law of partial pressures
gas mixture composition defined in terms of mole fraction \( (\chi_A = \text{mol A/mol total}) \)
relationship between mole fraction and partial pressure \( (\chi_A = P_A / P_{\text{tot}}) \)
calculations involving reactions that produce a gas when the gas is collected over water. i.e. \( P_{\text{tot}} = P_{\text{prod}} + P_{\text{H}_2\text{O}} \)
(\( \text{where } P_{\text{H}_2\text{O}} = \text{partial pressure of water vapor or vapor pressure of water} \))

- Kinetic Molecular Theory (KMT) of Gases
  theoretical model to explain observed gas behavior
calculation of average molecular speed or rms speed of a particle in a sample of gas using \( u = (3RT/M_n)^{1/2} \). Here,

\[
R = 8.314 \text{ kg•m}^2(\text{s}^3)\text{K}^{-1}•\text{mol}^{-1}
\]
and molar mass MUST be in units of kg/mol
relationship between molecular speed and \( T \), and molecular speed and molar mass
Maxwell's Distribution of Molecular Speeds – graphs and conclusions relating to speed and \( T \) and speed and molar mass

- Diffusion and Effusion - Graham's law of effusion; relationship between rate of effusion, time required for effusion, and molar mass

- Real Gases vs. Ideal Gases
  Intermolecular forces of attraction and repulsion within real gases cause them to behave non-ideally, especially at high \( P \) and low \( T \).
  Van der Waals equation can be used to calculate \( P \) of real gases, and how the \( P \) of a real gas compares to the \( P \) calculated when gases are assumed to behave ideally.

## Chapter 6: Thermochemistry

- Energy: kinetic vs. potential energy; energy transferred in the form of heat or work
  the change of internal energy of a system can be calculated: \( \Delta E = q + w \)
  \( PV \) work - gases are able to do work when they expand or are compressed: \( w = -P\Delta V \)
Units of energy: J, kJ, cal, kcal (and the conversions between them)
The First Law of Thermodynamics - the energy of the universe is constant
Terminology: system vs. surroundings; endo- vs. exothermic processes and the direction of energy flow
The convention of signs (+ or −) with thermodynamic data

- Enthalpy, \( \Delta H \): equate enthalpy change for a process to energy change for that process if it occurs at constant \( P \); \( \Delta H = q_P \)
  \( \Delta H = H_{\text{final}} - H_{\text{initial}} \)

\( \Delta H \) associated with physical changes: \( \Delta H_{\text{fus}}, \Delta H_{\text{fus}} = \Delta H_{\text{vap}} = \Delta H_{\text{sub}} \) where vap = vaporization, \( (g \leftrightarrow l) \),
  fus = fusion \( (l \leftrightarrow s) \), sub = sublimation \( (s \leftrightarrow g) \). This is not in your book!!

Remember that the \( \Delta H \) of the reverse of any change is \( = \) to the \( \Delta H \) of the forward change, but opposite in sign!

\( \Delta H \) associated with chemical changes: \( \Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}} \)
interpret amounts of energy released or absorbed by chemical reactions based on reaction stoichiometry
use \( \Delta H \) as a conversion factor between amount of reactant or product involved in a reaction and amount of heat absorbed or released.

\( \Delta H \) is +, the change is endothermic. If \( \Delta H \) is − the change is exothermic; what does this mean in terms of relative enthalpies of the initial and final states?
enthalpy diagrams.

- Calorimetry - an experimental technique to measure changes of energy related to chemical reactions.
  constant \( P \) calorimetry \( \rightarrow \) calculate \( \Delta H \)
  constant \( V \) calorimetry (bomb calorimetry) \( \rightarrow \) calculate \( \Delta E \)
  heat capacity and specific heat capacity
q = (specific heat)(mass)(ΔT)

\[
q_{\text{in}} = -q_{\text{out}}
\]
Keep track of the energy flow so that your signs make sense at the end of the calculation!
• Hess’s Law of Heat Summation

• Enthalpies of formation

  Standard States

  equations of formation – chemical equation for which $\Delta H$ is $\Delta H^\circ$ for compound

  Calculating $\Delta H^\circ_{\text{rxn}}$ from $\Delta H^\circ_\text{products}$ using: $\Delta H^\circ_{\text{rxn}} = \sum n \Delta H^\circ_\text{(products)} - \sum n \Delta H^\circ_\text{(reactants)}$

The following are some of the more important conversion factors that we have discussed in Ch. 4 - 6:
• Molar Concentration, or Molarity offers a relationship between the number of moles of solute dissolved per L of solution
• Molar Volume of a gas at STP offers a relationship between the mol of gas present and the volume occupied at STP (22.4 L/1 mol)
• $\Delta H$ associated with a balanced chemical equation offers a relationship between the quantity of heat released or absorbed and the quantity of reactant consumed or product formed

Some Review Problems:

1. Classify each of the following as a strong electrolyte, a weak electrolyte, or a nonelectrolyte:
   a. NaOH  
   b. NH$_4$Br  
   c. NH$_3$
   d. HNO$_3$  
   e. HClO$_4$  
   f. HBrO$_4$
   g. HC$_3$H$_7$O$_2$  
   h. K$_2$SO$_4$  
   i. BaSO$_4$
   j. AgCl  
   k. PbI$_2$  
   l. Pb(C$_2$H$_3$O$_2$)$_2$

2. Calculate the molarity of the following solutions:
   a. 0.0834 mol of Na$_2$SO$_4$ dissolved in enough water to form 650.0 mL of solution
   b. 1.45 g of NaCl dissolved in enough water to form 250.0 mL of solution

3. Calculate the mass (in g) of solute present in each of the following solutions:
   a. 1.00 x 10$^2$ mL of 0.150 M Na$_2$SO$_4$
   b. 2.50 x 10$^2$ mL of 0.0500 M KBrO$_3$

4. A sample of 0.0341 mol of iron (III) chloride was dissolved in water to give 25.0 mL of solution. Calculate the molarity of this solution. What is the molarity of chloride ions in this solution?

5. An experiment calls for 0.0353 g of potassium hydroxide, KOH. What volume (in mL) of 0.0176 M KOH is required?

6. You wish to prepare 0.12 M HNO$_3$ from a stock solution of nitric acid that is 15.8 M. What volume (in mL) of the stock solution is required to prepare 1.00 L of the 0.12 M HNO$_3$?

7. Calculate the volume of 0.250 M HNO$_3$ that will react completely with 42.4 mL of 0.150 M Na$_2$CO$_3$ according to the following reaction: 2 HNO$_3$ (aq) + Na$_2$CO$_3$ (aq) $\rightarrow$ 2 NaNO$_3$ (aq) + H$_2$O (l) + CO$_2$ (g)

8. Predict the products that will form when the following groups of reactants are mixed together. If a reaction occurs, write the balanced chemical equation, the full ionic equation, the net ionic equation for the reaction, and indicate whether it is a precipitation reaction or a neutralization reaction.
   a. Na$_2$CO$_3$ (aq) + BaCl$_2$ (aq) $\rightarrow$
   b. Na$_3$O (s) + 2HNO$_3$ (aq) $\rightarrow$
   c. Ca(HCO$_3$)$_2$ (s) + 2 HBr (aq) $\rightarrow$
   d. Silver nitrate + sodium carbonate $\rightarrow$ (both reactants are aqueous solutions)
   e. HC$_3$H$_7$O$_2$ (aq) + NaOH (aq) $\rightarrow$
   f. Sr(C$_2$H$_3$O$_2$)$_2$ (aq) + NiSO$_4$ (aq) $\rightarrow$
   g. Pb(NO$_3$)$_2$ (aq) + MgSO$_4$ (aq) $\rightarrow$
   h. sodium chloride + ammonium sulfate $\rightarrow$ (both reactants are aqueous solutions)
   i. NH$_3$ (aq) + HBr (aq) $\rightarrow$
   j. Calcium Hydroxide + Hydrochloric Acid $\rightarrow$ (both reactants are aqueous solutions)
   k. SO$_2$ (g) + H$_2$O (l) $\rightarrow$

9. Identify the Bronsted Acid and Bronsted Base in the following reactions:
   a. HSO$_3^-$ (aq) + H$_2$O (l) $\rightarrow$ SO$_3^{2-}$ (aq) + H$_3$O$^+$ (aq)
   b. NH$_4^+$ (aq) + CN$^-$ (aq) $\rightarrow$ NH$_3$ (aq) + HCN (aq)
10. Determine the oxidation number for the italicized element in each of the following species:
   a. $Ga_2O_3$
   b. $PbO_2$
   c. $KBrO_4$
   d. $ClF_4$
   e. $K_MnO_4$
   f. $NH_4^-$
   g. $IO^-$
   h. $C_2H_5O_3$
   i. $H_2PO_4^-$
   j. $NO_3^-$
   k. $CrO_3$
   l. $Na_2SO_3$

11. For each of the following reactions identify the element oxidized, the element reduced, the oxidizing agent, and the reducing agent.
   a. $P_4(s) + 5 O_2(g) \rightarrow P_5O_{10}(s)$
   b. $Co(s) + Cl_2(g) \rightarrow CoCl_2(s)$
   c. $2Al(s) + 3 F_2(g) \rightarrow 2 AlF_3(s)$
   d. $ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$

12. Balance the following redox equations using the method of balancing the half reactions.
   a. $Cr_2O_7^{2-}(aq) + C_2O_4^{2-}(aq) \rightarrow Cr^{3+}(aq) + CO_2(g)$ in acidic solution
   b. $MnO_4^-(aq) + HNO_2(aq) \rightarrow Mn^{2+}(aq) + NO_2^-(aq)$ in acidic solution
   c. $PbO_2(s) + Mn^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + MnO_2^-(aq)$ in acidic solution
   d. $HNO_2(aq) + Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq) + NO_3^-(aq)$ in acidic solution
   e. $S^{2-}(aq) + I_2(s) \rightarrow SO_4^{2-}(aq) + I^-(aq)$ in basic solution
   f. $MnO_2^-(aq) + NO_3^-(aq) \rightarrow MnO_2(s) + NO_3^-(aq)$ in basic solution
   g. $Pb(OH)_2^{2-}(aq) + ClO^-(aq) \rightarrow PbO_2(s) + Cl^-(aq)$ in basic solution
   h. $Zn(s) + NO_2^-(aq) \rightarrow Nh_3(aq) + Zn(OH)_2^{2-}(aq)$ in basic solution

13. A 3.33 g sample of iron ore is treated resulting in a solution of iron (II) sulfate. The resulting solution is then titrated with 0.150 M $K_2Cr_2O_7$ (see balanced equation given below). 41.4 mL of the potassium dichromate solution are required to reach the stoichiometric point in this titration. Calculate the mass percent iron in the ore sample.

14. Bone is dissolved in hydrochloric acid resulting in 50.0 mL of a solution containing calcium chloride, $CaCl_2$. To precipitate the calcium ion from this solution, an excess of potassium oxalate is added. 1.437 g of calcium oxalate is collected. Calculate the molar concentration of the $CaCl_2$ solution.

15. A solution of hydrochloric acid was prepared by measuring 10.00 mL of concentrated acid into a 1.000 L volumetric flask and adding water to the mark. Another solution was prepared by adding 0.530 g of anhydrous sodium carbonate to a 100.0 mL volumetric flask and adding water to the mark. Then, 25.00 mL of the sodium carbonate solution was pipetted into a flask and titrated with the dilute acid. The stoichiometric point was reached after 26.50 mL of the acid had been added. The balanced chemical equation that is taking place during the titration is:

$$Na_2CO_3(aq) + 2 HCl(aq) \rightarrow 2 NaCl(aq) + H_2CO_3(aq)$$

What is the molar concentration of the concentrated hydrochloric acid solution? (Hint: There are 2 problems here, a titration problem and a dilution calculation.)

16. Consider a 3.15 L sample of Ne gas at 21°C and a pressure of 0.951 atm. What would be the volume of this sample of Ne if the pressure were increased to 1.564 atm at constant temperature?

17. A sample of $N_2$ gas at 18°C and 760 mm Hg has a volume of 2.67 mL. What is the volume of this sample at STP?

18. In an experiment a heavy walled 5.00 L flask is filled with 7.68 g of methane gas, $CH_4$ at 19°C. Calculate the pressure exerted by the $CH_4$ (g).

19. Calculate the density (in g/L) of $NH_3$ at 22°C and 751 mm Hg.

20. Nitric acid is produced from nitric oxide, NO, which is in turn prepared from ammonia by the Ostwald process:

$$4 NH_3(g) + 5 O_2(g) \rightarrow 4 NO(g) + 6 H_2O(g)$$

What volume of oxygen gas at 35°C and 2.15 atm is needed to produce 50.0 g of NO?

21. Calculate the total pressure (in atm) exerted by a mixture of 0.0200 mols of He and 0.0100 mols of $H_2$ in a 5.00 L flask at 10°C.

22. An aqueous solution of ammonium nitrite decomposes when heated giving nitrogen gas:

$$NH_4NO_2(s) \rightarrow 2 H_2O(g) + N_2(g)$$

This reaction may be used to prepare pure $N_2$. What mass (in g) of ammonium nitrite was consumed if 4.16 dm³ of nitrogen gas was collected over water at 19°C and 97.8 kPa? The vapor pressure of water at 19°C is 16.5 mm Hg.
23. Calculate the rms speed of O₂ molecules 375 K.

24. Calculate the ratio of rates of effusion of H₂ and H₂S under identical experimental conditions.

25. 0.10 mol of I₂ vapor effuses from an opening in a vessel in 52 seconds. How long (in seconds) will it take 0.10 mol of H₂ to effuse under the same conditions?

26. A 2.30 g sample of white solid is vaporized in a 345 mL vessel. The pressure exerted by the vapor is 985 mm Hg at 148°C. Calculate the molar mass of the solid.

27. Sulfur is commonly found in coal in the form of iron pyrite, FeS₂. In the combustion of the coal, oxygen reacts with the iron pyrite to produce iron (III) oxide and sulfur dioxide:

\[ 4 \text{FeS}_2 (s) + 11 \text{O}_2 (g) \rightarrow 2 \text{Fe}_2\text{O}_3 (s) + 8 \text{SO}_2 (g) \]

a. Calculate the mass of Fe₂O₃ that is produced from the reaction of 75.0 L of O₂ at 23.3 atm and 150°C with an excess of iron pyrite.
b. The sulfur dioxide that is generated in (a) is dissolved in water to form 5.00 L of aqueous solution. What is the molar concentration of the resulting sulfurous acid, H₂SO₃ solution?

28. Small quantities of hydrogen gas in the laboratory can be generated by the reaction of dilute hydrochloric acid with zinc metal. When 0.40 g of impure zinc was reacted with an excess of hydrochloric acid, 127 mL of hydrogen was collected over water at 17°C and a total pressure of 737.7 mmHg. The vapor pressure of water at 17°C is 14.5 mm Hg

a. What amount (in moles) of H₂ was collected?
b. What is the mass percent of zinc in the impure zinc sample?

29. A 15 mL sample of ammonia gas at 100. Torr and 30°C is mixed with 25 mL of hydrogen chloride at 25°C and 150 Torr. The reaction that occurs is: \( \text{NH}_3 (g) + \text{HCl} (g) \rightarrow \text{NH}_4\text{Cl} (s) \)

a. Calculate the mass of NH₄Cl that forms.
b. Identify the gas in excess. After the reaction is complete, what pressure will the excess gas exert in a 40.0 mL container at 27°C?

30. Ammonium sulfate, an important fertilizer, can be prepared by the reaction of ammonia with sulfuric acid:

\[ 2 \text{NH}_3 (g) + \text{H}_2\text{SO}_4 (aq) \rightarrow (\text{NH}_4)_2\text{SO}_4 (aq) \]

Calculate the volume of NH₃ needed at 20°C and 25.0 atm to react with 1.50 x 10² kg of H₂SO₄.

31. A gas of unknown molar mass was allowed to effuse through a small opening under constant pressure conditions. It required 72 sec for 1 L of the gas to effuse. Under identical experimental conditions it required 28 sec for 1 L of O₂ gas to effuse. Calculate the molar mass of the unknown gas.

32. When 2 mols of potassium chlorate crystals decompose to potassium chloride crystals and oxygen gas at constant temperature and pressure, 44.7 kJ of heat is given off. Write the thermochemical equation corresponding to this reaction.

33. Hydrogen, H₂, is used as a rocket fuel. The hydrogen is burned in oxygen to produce water vapor. What is the enthalpy change per gram of H₂ (you are solving for kJ/g H₂)? 2 \( \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O} (g) \quad \Delta H = -484 \text{ kJ} \)

34. Ammonia burns in the presence of a copper catalyst to form nitrogen gas. What is the enthalpy change associated with burning 25.6 g of ammonia?

\[ 4 \text{NH}_3 (g) + 3 \text{O}_2 (g) \rightarrow 2 \text{N}_2 (g) + 6 \text{H}_2\text{O} (g) \quad \Delta H = -1267 \text{ kJ} \]

35. Calculate the amount of heat evolved (in kJ) when 13.4 L of SO₂ at STP and 15.0 L of O₂ at 18°C at 1.02 atm are combined. The balanced chemical equation for this reaction follows:

\[ 2 \text{SO}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{SO}_3 (g) \quad \Delta H = -198 \text{ kJ} \]

36. Hydrogen cyanide is a highly poisonous, volatile liquid. It can be prepared by the reaction:

\[ \text{CH}_4 (g) + \text{NH}_3 (g) \rightarrow \text{HCN} (g) + 3 \text{H}_2 (g) \]

What is \( \Delta H \) of this reaction at constant pressure? Use the following thermochemical equations:

\[ \text{N}_2 (g) + 3 \text{H}_2 (g) \rightarrow 2 \text{NH}_3 (g) \quad \Delta H = -91.8 \text{ kJ} \]
\[ \text{C} \text{ (graphite)} + 2 \text{H}_2 (g) \rightarrow \text{CH}_4 (g) \quad \Delta H = -74.9 \text{ kJ} \]
\[ \text{H}_2 (g) + 2 \text{C} \text{ (graphite)} + \text{N}_2 (g) \rightarrow 2 \text{HCN} (g) \quad \Delta H = 270.3 \text{ kJ} \]

37. The first step in the preparation of lead from its ore (galena, PbS) consists of roasting the ore:

\[ 2 \text{PbS} (s) + 3 \text{O}_2 (g) \rightarrow 2 \text{SO}_2 (g) + 2 \text{PbO} (s) \]

Calculate the standard enthalpy change for this reaction, using enthalpies of formation.
38. The Group IIA carbonates decompose when heated. For example: \( \text{BaCO}_3 (s) \rightarrow \text{BaO} (s) + \text{CO}_2 (g) \). Use enthalpies of formation to calculate the heat required to decompose 10.0 g of barium carbonate. For \( \text{BaO} \), \( \Delta H_f^o = -582 \text{ kJ/mol} \).

39. Calculate the enthalpy change for the vaporization of 100.0 g \( H_2O \) at 373.2 K. The \( \Delta H_{vap} \) of \( H_2O (l) \) = 40.7 kJ/mol.

40. Calculate \( \Delta H \) of melting of 600 g (3 s.f.) of solid \( \text{NH}_3 \) at its freezing point, 195.3 K. \( \Delta H_{melt} = 5.65 \text{ kJ/mol} \).

41. Calculate the enthalpy of reaction for the synthesis of hydrogen chloride gas from the following data:

\[
\begin{align*}
\text{H}_2 (g) + \text{Cl}_2 (g) & \rightarrow 2 \text{HCl} (g) \text{ (target eqn)} \\
\text{NH}_3 (g) + \text{HCl} (g) & \rightarrow \text{NH}_4\text{Cl} (s) \quad \Delta H = -176.0 \text{ kJ} \\
\text{N}_2 (g) + 3 \text{H}_2 (g) & \rightarrow 2 \text{NH}_3 (g) \quad \Delta H = -92.22 \text{ kJ} \\
\text{N}_2 (g) + 4 \text{H}_2 (g) + \text{Cl}_2 (g) & \rightarrow 2 \text{NH}_4\text{Cl} (s) \quad \Delta H = -628.86 \text{ kJ}
\end{align*}
\]

42. Write chemical equations for the formation of the standard states of the following compounds:

a. \( \text{KClO}_3 (s) \),  

b. \( \text{H}_3\text{NCH}_2\text{COOH} (s) \),  

c. \( \text{Al}_2\text{O}_3 (s) \)

43. Calculate the standard enthalpy of formation of \( \text{PCl}_3 (s) \). \( \Delta H_f^o \) of \( \text{PCl}_3 (l) \) = 137.33 kJ/mol

\( \text{PCl}_3 (l) + \text{Cl}_2 (g) \rightarrow \text{PCl}_5 (s) \quad \Delta H_f^o = -124 \text{ kJ} \)

44. How much heat (in kJ) is required to convert 100. g of solid benzene (\( \text{C}_6\text{H}_6 \)) at \(-5^\circ C\) to gaseous benzene at \(150^\circ C\)? For benzene: freezing point = 5.7°C, boiling point = 80.3°C, \( \Delta H_{fus} = 9.87 \text{ kJ/mol} \), \( \Delta H_{vap} = 30.8 \text{ kJ/mol} \), specific heat of \( \text{C}_6\text{H}_6 (s) = 0.556 \text{ J/g°C} \), specific heat of \( \text{C}_6\text{H}_6 (l) = 1.050 \text{ J/g°C} \), specific heat of \( \text{C}_6\text{H}_6 (g) = 0.743 \text{ J/g°C} \).

Chem 171

Review Exam 2: Answers

1. a. strong  
   d. strong  
   g. weak  
   j. non

b. strong  
   e. strong  
   h. strong  
   k. non

   c. weak  
   f. weak  
   i. non  
   l. strong

2. a. .128 M  
   b. .0992 M

3. a. 2.13 g  
   b. 2.09 g

4. 1.36 M \( \text{FeCl}_3 \) (aq), 4.08 M \( \text{Cl}^- \) (aq)

5. 35.7 mL

6. 7.6 mL

7. 50.9 mL

8. a. ppt’n reaction

\( \text{Na}_2\text{CO}_3 (aq) + \text{BaCl}_2 (aq) \rightarrow \text{BaCO}_3 (s) + 2 \text{NaCl} (aq) \)

\begin{align*}
2 \text{Na}^+ (aq) + \text{CO}_3^{2-} (aq) + \text{Ba}^{2+} (aq) + 2 \text{Cl}^- (aq) & \rightarrow \text{BaCO}_3 (s) + 2 \text{Na}^+ (aq) + 2 \text{Cl}^- (aq) \\
\text{Ba}^{2+} (aq) + \text{CO}_3^{2-} (aq) & \rightarrow \text{BaCO}_3 (s)
\end{align*}

b. acid/base - oxides of metals are basic

\( \text{Na}_2\text{O} (s) + 2 \text{HNO}_3 (aq) \rightarrow 2 \text{NaNO}_3 (aq) + \text{H}_2\text{O} (l) \)

\( \text{Na}_2\text{O} (s) + 2 \text{H}^+ (aq) + 2 \text{NO}_3^- (aq) \rightarrow 2 \text{Na}^+ (aq) + 2 \text{NO}_3^- (aq) + 2 \text{H}_2\text{O} (l) \)

\( \text{Na}_2\text{O} (s) + 2 \text{H}^+ (aq) \rightarrow 2 \text{Na}^+ (aq) + 2 \text{H}_2\text{O} (l) \)

c. acid/base – hydrogen carbonates are basic

\( \text{Ca} (\text{HCO}_3)_2 (s) + 2 \text{HBr} (aq) \rightarrow \text{CaBr}_2 (aq) + 2 \text{H}_2\text{O} (l) + 2 \text{CO}_2 (g) \)

\( \text{Ca} (\text{HCO}_3)_2 (s) + 2 \text{H}^+ (aq) + 2 \text{Br}^- (aq) \rightarrow \text{Ca}^{2+} (aq) + 2 \text{Br}^- (aq) + 2 \text{H}_2\text{O} (l) + 2 \text{CO}_2 (g) \)

\( \text{Ca} (\text{HCO}_3)_2 (s) + 2 \text{H}^+ (aq) \rightarrow \text{Ca}^{2+} (aq) + 2 \text{H}_2\text{O} (l) + 2 \text{CO}_2 (g) \)
d. ppt'n reaction
2 AgNO₃ (aq) + Na₂CO₃ (aq) $\rightarrow$ Ag₂CO₃ (s) + 2 NaNO₃ (aq)
2 Ag⁺ (aq) + 2 NO₃⁻ (aq) + 2 Na⁺ (aq) + CO₃²⁻ (aq) $\rightarrow$ Ag₂CO₃ (s) + 2 Na⁺ (aq) + 2 NO₃⁻ (aq)
2 Ag⁺ (aq) + CO₃²⁻ (aq) $\rightarrow$ Ag₂CO₃ (s)

e. acid/base
HC₃H₄O₂ (aq) + NaOH (aq) $\rightarrow$ NaC₃H₄O₂ (aq) + H₂O (l)
HC₃H₄O₂ (aq) + Na⁺ (aq) + OH⁻ (aq) $\rightarrow$ Na⁺ (aq) + C₃H₄O₂⁻ (aq) + H₂O (l)
HC₃H₄O₂ (aq) + OH⁻ (aq) $\rightarrow$ C₃H₄O₂⁻ (aq) + H₂O (l)

f. ppt'n reaction
Sr(C₂H₃O₂)₂ (aq) + NiSO₄ (aq) $\rightarrow$ SrSO₄ (s) + Ni(C₂H₃O₂)₂ (aq)
Sr²⁺ (aq) + 2 C₂H₃O₂⁻ (aq) + Ni²⁺ (aq) + SO₄²⁻ (aq) $\rightarrow$ SrSO₄ (s) + Ni²⁺ (aq) + 2 C₂H₃O₂⁻ (aq)
Sr²⁺ (aq) + SO₄²⁻ (aq) $\rightarrow$ SrSO₄ (s)

g. ppt'n reaction
Pb(NO₃)₂ (aq) + MgSO₄ (aq) $\rightarrow$ PbSO₄ (s) + Mg(NO₃)₂ (aq)
Pb²⁺ (aq) + 2 NO₃⁻ (aq) + Mg²⁺ (aq) + SO₄²⁻ (aq) $\rightarrow$ PbSO₄ (s) + Mg²⁺ (aq) + 2 NO₃⁻ (aq)
Pb²⁺ (aq) + SO₄²⁻ (aq) $\rightarrow$ PbSO₄ (s)

h. no reaction

i. acid/base
NH₃ (aq) + HBr (aq) $\rightarrow$ NH₄Br (aq)
NH₃ (aq) + H⁺ (aq) + Br⁻ (aq) $\rightarrow$ NH₄⁺ (aq) + Br⁻ (aq)
NH₃ (aq) + H⁺ (aq) $\rightarrow$ NH₄⁺ (aq)

j. acid/base
Ca(OH)₂ (aq) + 2 HCl (aq) $\rightarrow$ CaCl₂ (aq) + 2 H₂O (l)
Ca²⁺ (aq) + 2 OH⁻ (aq) + 2 H⁺ (aq) + 2 Cl⁻ (aq) $\rightarrow$ CaCl₂ (aq) + 2 Cl⁻ (aq) + 2 H₂O (l)
2 H⁺ (aq) + 2 OH⁻ (aq) $\rightarrow$ 2 H₂O (l)

k. acid/base - oxides of non-metals are acidic
SO₃ (g) + H₂O (l) $\rightarrow$ H₂SO₄ (aq)
SO₃ (g) + H₂O (l) $\rightarrow$ H⁺ (aq) + HSO₄⁻ (aq) \textit{OR} $\rightarrow$ 2 H⁺ (aq) + SO₄²⁻ (aq)

9. a. HSO₃⁻ acid, H₂O base
   b. NH₄⁺ acid, CN⁻ base

10. a. +3 b. +4 c. +7 d. +3
    e. +6 f. -3 g. +1 h. 0
    i. +5 j. +5 k. +6 l. +4

11. a. P oxidized, O reduced, P₄ reducing agent, O₂ oxidizing agent
    b. Co oxidized, Cl reduced, Co reducing agent, Cl₂ oxidizing agent
    c. Al oxidized, F reduced, Al reducing agent, F₂ oxidizing agent
    d. C oxidized, Zn reduced, C reducing agent, ZnO oxidizing agent

12. a. 3 C₂O₄²⁻ (aq) + 14 H⁺ (aq) + Cr₂O₇⁻ (aq) $\rightarrow$ 6 CO₂ (g) + 2 Cr³⁺ (aq) + 7 H₂O (l)
b. HNO₂ (aq) + H⁺ (aq) + MnO₂ (s) $\rightarrow$ Mn³⁺ (aq) + H₂O (l) + NO⁻ (aq)
c. 4 H⁺ (aq) + 5 PbO₂ (s) + 5 SO₄²⁻ (aq) + 2 Mn²⁺ (aq) $\rightarrow$ 5 PbSO₄ (s) + 2 H₂O (l) + 2 MnO₄⁻ (aq)
d. 3 HNO₂ (aq) + 5 H⁺ (aq) + Cr₂O₇⁻ (aq) $\rightarrow$ 3 NO⁻ (aq) + 2 Cr³⁺ (aq) + 4 H₂O (l)
e. 8 OH⁻ (aq) + S²⁻ (aq) + 4 I₂ (s) $\rightarrow$ 8 I⁻ (aq) + SO₄²⁻ (aq) + 4 H₂O (l)
f. 2 OH⁻ (aq) + 3 NO₂⁻ (g) + MnO₂⁻ (aq) $\rightarrow$ 3 NO₃⁻ (aq) + MnO₂ (s) + H₂O (l)
g. Pb(OH)₄²⁻ (aq) + Cl⁻ (aq) $\rightarrow$ Pb₂O₃ (s) + Cl⁻ (aq) + H₂O (l) + 2 OH⁻ (aq)
h. 7 OH⁻ (aq) + 4 Zn (s) + 6 H₂O (l) + NO₃⁻ (aq) $\rightarrow$ 4 Zn(OH)₄²⁻ (aq) + NH₃ (aq)
13. 62.5 
14. 0.224 M
15. 9.43 M
16. 1.92 L
17. 0.00251 L
18. 2.30 atm
19. 0.695 g/L
20. 24.5 L
21. 0.139 atm
22. 10.5 g
23. 541 m/s
24. 4.11
25. 4.6 s
26. 177 g/mol
27. a. 146 g Fe₂O₃
   b. .732 M
28. a. .00508 mol
   b. 83% pure
29. a. 0.00426 g
   b. 0.0745 atm
30. 3.0 x 10³ L
31. 212 g/mol
32. 2 KClO₃ (s) → 2 KCl (s) + 3 O₂ (g)  ΔH = −44.7 kJ
33. −120.0 kJ
34. −476 kJ
35. −59.2 kJ
36. +256 kJ
37. −835.8 kJ
38. +12.2 kJ
39. +226 kJ
40. +199 kJ
41. −184.86 kJ
42. a. K(s) + 1/2 Cl₂ (g) + 3/2 O₂ (g) → KClO₃ (s)
   b. 5/2 H₂ (g) + 1/2 N₂ (g) + 2 C (gr.) + O₂ (g) → H₂NCH₂COOH (s)
   c. 2 Al (s) + 3/2 O₂ (g) → Al₂O₃ (s)
43. +13.33 kJ
44. +65.6 kJ