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Honors Chemistry Exam Review

Essential Standard 1.1: Analyze the Structure of Atoms & Ions

1. Fill in the chart below.

Subatomic Particle	Location	Relative charge	Mass
Proton			
Neutron			
Electron			

2. Identify the following elements.

   a. $^{235}_{92}X$: _____ b. $^{17}_{8}X$: _____

3. Write the isotopic symbol for the following:
   a. An element that has 17 protons, 18 electrons, 18 neutrons _____

   b. An element that has a 20 protons, 18 electrons, and 21 neutrons _____

   c. An element that has 93 protons, 93 electrons, and 154 neutrons _____

4. Calculate the atomic mass of an element that has 3 isotopes with the following mass and relative abundance data.

Isotope	Mass	Relative abundance
1	45.699	33.26%
2	46.799	44.22%
3	47.899	22.52%

5. Using the Bohr Diagram in your reference packet, answer the following questions.
   a. What is the wavelength of light emitted when an electron moves from n=3 to n=1?

   b. What is the type of light emitted?

   c. Calculate the frequency of the emitted light.

   d. Calculate the energy of the emitted light.

6. Write complete electron configurations for the following:

   a. O _____ b. Fe _____

   c. Ca^{+2}_____ d. P^{-3}_____

7. Write noble gas configurations for the following:

   a. Cs _____ b. Br _____

8. Define the following terms:
   a. Quanta

   b. Energy

   c. Wavelength

   d. Frequency

   e. Energy level

   f. Photons

9. When electrons _____ energy they become _____ and move to a higher energy level.

10. When electrons _____ energy they become _____ and move to a lower energy level.

11. The relationship between wavelength and frequency is (inverse/direct).

12. The relationship between energy and frequency is (inverse/direct).

13. Which of the following statements was NOT made by Neils Bohr in description of the atom?
    a. An electron circles the nucleus in fixed energy ranges called orbits.
    b. An electron can neither gain or lose energy inside this orbit, but could move up or down to another orbit
    c. The lowest energy orbit is closest to the nucleus
    d. Energy levels are divided into sublevels.

14. Fill in the table below

Particle	Symbol	Mass	Charge	Penetrating ability
Alpha				
Beta				
Gamma				

15. Complete the following reactions.

a. $^{234}_{92}U \rightarrow {}^{4}_{2}He + \_\_\_\_\_$

b. $^{18}_{8}O + {}^{0}_{-1}e \rightarrow \_\_\_\_\_ + {}^{1}_{0}n$

c. $^{1}_{1}H + {}^{2}_{1}H \rightarrow \_\_\_\_\_$

d. $^{14}_{6}C \rightarrow {}^{0}_{-1}e + \_\_\_\_\_$

16. Write balanced nuclear equations for the following.

a. Alpha bombardment of Thorium-290.

b. Beta decay of Radium-222

17. Determine the half life of a radioactive isotope that decays from 100.0mg to 44.3mg in 24.0 hours.

18. How much of a 25.0g sample of $^{14}_{6}C$ remains after 100,000 years? The half life of $^{14}_{6}C$ is 5730 years?

19. How many grams were originally present in a sample that decays to 5.0g in 55.3 hours if the half life is 2.0 days?

20. Identify the following as either fission or fusion.

a. Occurs in the stars like the Sun _____

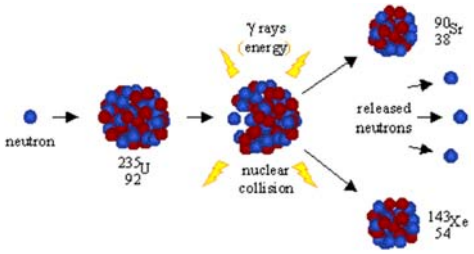
b. Used to generate energy we use in our homes. _____

c. Combining two small nuclei to form a larger nucleus. _____

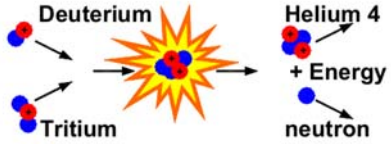
d. Creates more energy. _____

e. Requires more energy. _____

f. Splits a large nucleus into smaller nuclei. _____

g. 

_____

h. 

_____

Constructed Response Examples.

1. Write your answers on a separate sheet of paper.

2. Be sure to write your name on each page.

1. An element has two isotopes if 7.42% exists as ^{6}X (6.015amu) and 92.58% exists as ^{7}X (7.016amu).
   - Explain the difference between the mass of an isotope, the mass number of an isotope, and the atomic mass of an element.
   - What is the average atomic mass of element X.
   - Identify element X.
   - Write the isotopic symbol for the most abundant isotope of element X.
2. We are exposed to radiation every day because unstable isotopes undergo decay constantly.
   - What are two types of radioactive decay which result in the creation of a new element?
   - What are the symbols used for these two types of decay?
   - What is the resulting nuclide when Uranium-234 undergoes alpha decay?

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Honors Chemistry Exam Review

Essential Standard 1.2: Understand the bonding that occurs in simple compounds in terms of bond type, strength, and properties.

1. Identify the following statements about bonding as ionic, covalent, or metallic.
   a. Sea of electrons _____
   b. Formed by a metal and a nonmetal _____
   c. Formed by nonmetals only _____
   d. Conducts electricity as a solid _____
   e. Poor conductor of electricity _____
   f. Can be solid, liquid, or gas at room temperature _____
   g. Forms crystalline solids _____
   h. Have high melting points _____
   i. Have low boiling points _____
   j. Have high electrical conductivity in molten state _____
   k. Have high electrical conductivity in aqueous solution _____
   l. Are ductile _____
   m. Have low melting points _____
   n. Have luster _____
   o. Forms through a transfer of electrons _____
   p. Forms when atoms share electrons _____

2. A positively charged ion is called a(n) _____. A negatively charged atom is called a(n) _____.

3. If an element has 6 valence electrons what charge will it likely form?

4. If an element has 2 valence electrons what charge will it likely form?

5. Why do atoms form ions?
6. Draw the lewis structure, identify the geometry, and identify the polarity for the following:
   a. Phosphorus trichloride

   b. Boron trihydride

c. Carbon dioxide

d. Sulfur dioxide

e. Carbon tetrafluoride

7. List the 7 diatomic elements.

8. Write the name for the following:

a. $NaBr$ _____

b. $Ca(NO_3)_2$ _____

c. Li_2SO_4 _____

d. $FeBr_2$ _____

e. $Be(OH)_2$ _____

f. SnO_2 _____

g. N_2S _____

h. PH_3 _____

i. P_2Br_4 _____

j. $HClO_3$ _____

k. H_2SO_3 _____

l. HBr _____

9. Write the formula for the following:
    a. potassium iodide _____
    b. magnesium acetate _____
    c. aluminum chloride _____
    d. nickel (III) nitrate _____
    e. calcium carbonate _____
    f. lead (IV) sulfate _____
    g. beryllium phosphide _____
    h. Iron (III) carbide _____
    i. dinitrogen trioxide _____
    j. phosphorus pentafluoride _____
    k. sulfur dibromide _____
    l. diboron tetrahydride _____
    m. acetic acid _____
    n. nitric acid _____
    o. hydroiodic acid _____

10. Explain what the phrase "like dissolved like" means.

11. Which type of bond (single, double, or triple) has the most energy (is harder to break)?

12. Define the following terms:
    a. Intermolecular forces

    b. Hydrogen bonding

    c. Dipole-dipole force

    d. London dispersion forces

13. Which of the intermolecular forces is strongest? Weakest?

Constructed Response Examples.

1. *Write your answers on a separate sheet of paper.*
2. *Be sure to write your name on each page.*

1. Polyatomic ions are ion which consist of many atoms and are common in ionic compounds.
    - Draw the lewis structures for the following polyatomic ions: nitrate, sulfate, carbonate, and ammonium
    - Identify the geometry and polarity for each of the ions above.
    - Write the formula and name for the compound that forms between an sodium and nitrate.
    - Write the formula and name for the compound that forms between an calcium and sulfate.
    - Write the formula and name for the compound that forms between a iron with a charge of +2 and carbonate.
    - Write the formula and name for the compound that forms between ammonium and bromine.

2. There are 3 different types of bonding: ionic, covalent, and metallic.
    - Describe how each type of bond forms.
    - List 3 properties for each type of compound which results from each bond type.
    - Which of the 3 types of bonds is the weakest and what is the relationship between bond strength and the melting point?
    - Why are intermolecular forces weaker than ionic, covalent, or metallic bonds?

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Honors Chemistry Exam Review

Essential Standard 1.3: Understand the physical and chemical properties of atoms based on their position on the Periodic Table.

1. The rows on the periodic table are called _____. The columns on the periodic table are called _____.

2. Fill in the table below:

Group	**Name**	**Valence Electrons**	**Charge (oxidation state)**
IA			
IIA			
VIIA			
VIIIA			
B	Transition Metals	Varies (most have at least 2)	varies

3. Reactivity of metals (decreases/increases) down the group, but reactivity for nonmetals (decreases/increases) down the group. Therefore the most active metal is _____ and the most active nonmetal is _____.

4. Metals are on the _____ side of the periodic table. Nonmetals are on the _____ side of the periodic table. Metalloids are along the _____.

5. Classify the following elements as either a metal, nonmetal, or metalloid.

   a. Fe _____ b. Si _____ c. Ar _____

   d. Ca _____ e. U _____ f. O _____

6. Define the following terms:
   a. atomic radius

   b. Ionic radius

   c. Electronegativity

   d. Ionization energy

   e. Electron affinity

7. The atomic radius (increases/decreases) down a group and (increases/decreases) across a period.

8. Cations are (larger/smaller) than their respective neutral atom. Anions are (larger/smaller) than their respective neutral atom.

9. The ionization energy (increases/decreases) down a group and (increases/decreases) across a period.

10. The electronegativity (increases/decreases) down a group and (increases/decreases) across a period.

11. Write the orbital notation for Bromine.

12. How many valence electrons do the following have?

a. $1s^2 2s^2 2p^6 3s^1$ _____ b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ _____

c. $1s^2 2s^2 2p^6 3s^2 3p^3$ _____ d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d1^0 4p^5$ _____

13. How many electrons would you expect the following to lose or gain. (for example gain 2, or lose 1)

a. $1s^2 2s^2 2p^6 3s^1$ _____ b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ _____

c. $1s^2 2s^2 2p^6 3s^2 3p^3$ _____ d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ _____

14. What would be the charge (oxidation state) for the following?

a. $1s^2 2s^2 2p^6 3s^1$ _____ b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ _____

c. $1s^2 2s^2 2p^6 3s^2 3p^3$ _____ d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d1^0 4p^5$ _____

15. Put the following groups of elements in increasing order of ionization energy

a. C, Fe, Fr _____

b. Xe, Co, Na _____

16. Put the following groups of elements in decreasing order of electronegativity.

a. C, Fe, Fr _____

b. Xe, Co, Na _____

Constructed Response Examples.

1. *Write your answers on a separate sheet of paper.*
2. *Be sure to write your name on each page.*

1. Atomic size is one of many trends of the periodic table.
    - Describe one reason atomic size may vary among the elements of the periodic table.
    - List the correct order of aluminum, magnesium, phosphorus, silicon, sodium, and sulfur based on decreasing atomic size.
    - Explain the relationship between atomic size and ionization energy.

2. Metal, nonmetals, and metalloids have different properties.
    - List 3 properties of each.
    - How does the reactivity of metals differ from the reactivity of nonmetals based on their location on the periodic table?
    - Why are alkali metals more reactive than alkaline earth metals?
    - Why are halogens more reactive than noble gases?

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Honors Chemistry Exam Review

Essential Standard 2.1: Understand the relationship among pressure, temperature, volume, and phase.

1. Define the following terms:
    a. Temperature

    b. Heat

    c. Kinetic energy

    d. Potential energy

    e. Joule

    f. Calorie

    g. Celsius

    h. Kelvin

    i. Melting

    j. Boiling

    k. Freezing

    l. Condensation

    m. Sublimation

    n. Deposition

    o. Endothermic

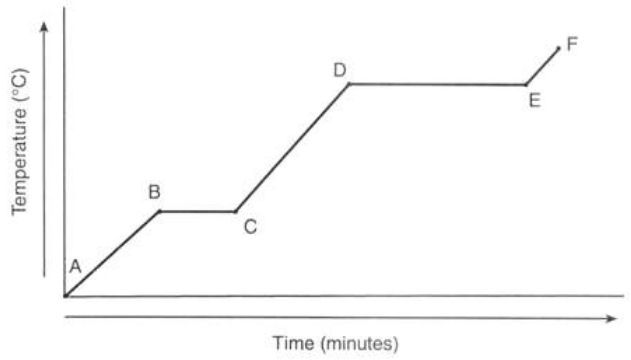
    p. Exothermic

    q. Specific heat

    r. Heat of fusion

    s. Heat of vaporization

2. Identify the following on the heating curve below:



Solid: _____

melting: _____

Liquid: _____

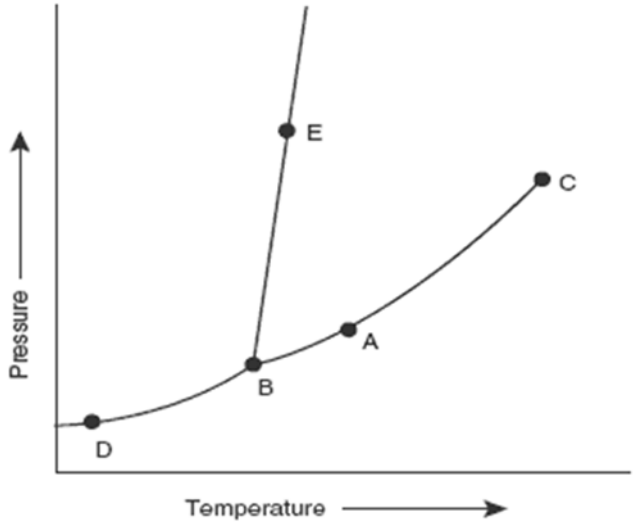
Condensation: _____

Vapor: _____

Which regions represent changes in kinetic energy?

Which regions represent changes in potential energy?

3. Identify the following on the phase diagram below



Sublimation: _____

Melting: _____

Boiling: _____

Triple point: _____

Critical point: _____

Label the diagram to show where solid (S), liquid (L), and gas (G) phases are located.

What happens to the substance if the pressure is increased at a low temperature?

What happens to the substance if the temperature in increased at a high pressure?

3. In a closed system, heat is neither _____ or _____ only _____ between components of the system.

4. Calculate the mass of aluminum that would increase its temperature from 30.0°C to 40.0°C when 2500J of energy are absorbed.

5. How many grams of ice can be melted if 3500J of energy are absorbed?

6. How many joules of energy are released when 150.0g of water vapor are condensed to liquid water?

7. Calculate the specific heat of a substance if 25.0g of the substance absorbs 3400J of energy and increases its temperature from 10.0°C to 25.0°C

8. A 97 g sample of gold at 785°C is dropped into 323 g of water, which has an initial temperature of 15°C. If gold has a specific heat of 0.129 J/g·°C, what is the final temperature of the mixture?

9. Identify the following gas law equations:

   a. $PV = nRT$ _____

   b. $P_1V_1 = P_2V_2$ _____

   c. $P_1/T_1 = P_2/T_2$ _____

   d. $V_1/T_1 = V_2/T_2$ _____

   e. $P_1V_1/T_1 = P_2V_2/T_2$ _____

   f. $P_T = P_1+P_2+P_3+...$ _____

9. How many liters do 5.50mol of oxygen occupy at STP?

10. How many moles of argon occupy 3.40L at 1.2atm and 25.0°C?

11. If I have 5.6 liters of gas in a piston at a pressure of 1.5 atm and compress the gas until its volume is 4.8 L, what will the new pressure inside the piston be?

12. Calcium carbonate decomposes at 1200^0 C to form carbon dioxide and calcium oxide. If 25 liters of carbon dioxide are collected at 1200^0 C, what will the volume of this gas be after it cools to 25^0 C?

13. A toy balloon has an internal pressure of 1.05 atm and a volume of 5.0 L. If the temperature where the balloon is released is 20^0 C, what will happen to the volume when the balloon rises to an altitude where the pressure is 0.65 atm and the temperature is -15^0 C?

14. Two flasks are connected with a stopcock. The first flask has a volume of 5 liters and contains nitrogen gas at a pressure of 0.75 atm. The second flask has a volume of 8 L and contains oxygen gas at a pressure of 1.25 atm. When the stopcock between the flasks is opened and the gases are free to mix, what will the pressure be in the resulting mixture?

Constructed Response Examples.

1. *Write your answers on a separate sheet of paper.*
2. *Be sure to write your name on each page.*

1. During a laboratory experiment, 75 grams of water at 100C is transformed into steam at 100C.
    a. Describe how this experiment confirms the law of conservation of energy.
    b. How much heat energy is needed to completely change the state of the water? Show your work.

2. The following data was collected during a laboratory experiment.

	Trial 1	**Trial 2**
Mass of metal	15.0g	15.0g
Mass of water	100.0g	100.0g
Initial temperature of water	25.0°C	25.0°C
Initial temperature of metal	100.0°C	100.0°C
Final temperature of water	25.35°C	25.25°C
Final temperature of metal		

    a. Identify the metal. Show your work.
    b. Describe how this experiment confirms the law of conservation of energy.

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Honors Chemistry Exam Review

Essential Standard 2.2: Analyze chemical reactions in terms of quantities, product formation, and energy

1. In order for molecules to react they must _____ with enough

   _____ and in the correct _____.

Use the diagrams below to answer questions 2-5

Diagram A

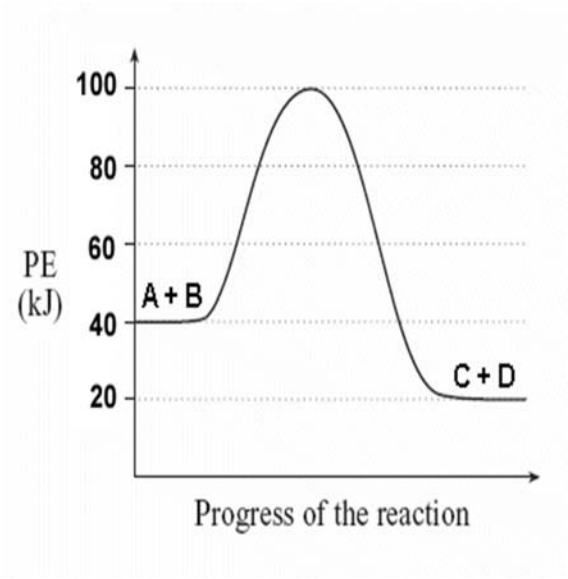
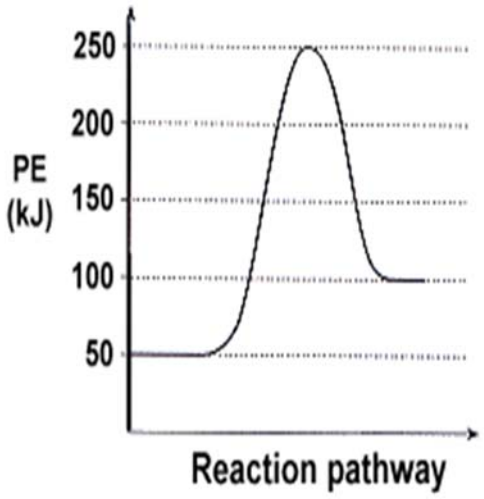


Diagram B



2. Which diagram is endothermic? _____ exothermic? _____

3. In diagram A, what is the energy of the activated complex? _____

4. In diagram B, what is the energy of the reaction (ΔH)? _____

5. In diagram A, what is the activation energy? _____

6. The sign of ΔH is _____ for endothermic reactions and _____ for exothermic reactions.

7. List 5 indicators that a chemical reaction has occurred.

8. Identify the type of reaction, predict the products, and balance the following:

   a. _____ Na + _____ O_2 →

      Type of reaction: _____

   b. _____ NH_3 →

      Type of reaction: _____

   c. _____ C_5H_9O + _____ O_2 →

      Type of reaction: _____

   d. _____ Pb + _____ H_3PO_4 →

      Type of reaction: _____

   e. _____ Li_3N + _____ NH_4NO_3 →

      Type of reaction: _____

   f. _____ HBr + ___ $Al(OH)_3$ →

      Type of reaction: _____

9. $LiOH + HBr \rightarrow LiBr + H_2O$

   If you start with 10.0 grams of lithium hydroxide, how many grams of lithium bromide will be produced?

10. $C_2H_4 + 3\ O_2 \rightarrow 2\ CO_2 + 2\ H_2O$

    If you start with 4.5×10^{22} molecules of ethylene (C_2H_4), how many liters of carbon dioxide will be produced at STP?

11. $Mg + 2\ NaF \rightarrow MgF_2 + 2\ Na$

    If you start with 5.5 grams of sodium fluoride, how many grams of magnesium fluoride will be produced?

12. $2\ HCl + Na_2SO_4 \rightarrow 2\ NaCl + H_2SO_4$

    If you start with 20 grams of hydrochloric acid, how many molecules of sulfuric acid will be produced?

13. What is the empirical formula for a compound which contains 0.0134 g of iron, 0.00769 g of sulfur and 0.0115 g of oxygen?

14. Find the empirical formula for a compound which contains 32.8% chromium and 67.2% chlorine.

15. Find the molecular formula of a compound with an empirical formula of C_2OH_4 and a molar mass of 88 grams per mole.

16. What is the percent composition of potassium carbonate?

17. A 15.00 gram sample of a hydrate was found to contain 7.05 grams of water. If the anhydrous salt left was sodium sulfate, determine the formula of the hydrate.

Constructed Response Examples.

1. *Write your answers on a separate sheet of paper.*
2. *Be sure to write your name on each page.*

1. We can write a chemical reaction for any combination of atoms and/or ions, but that does not mean that reaction occurs.
    - For the following reactions: write a balanced chemical reaction:
        a. Aluminum reacts with sodium nitrate.
        b. Potassium chloride reacts with lead (II) nitrate.
        c. Zinc reacts with hydrochloric acid.
        d. Magnesium acetate reacts with copper (II) nitrate.
    - For each of the reactions (a-d) determine if the reaction occurs. For single replacement reactions, use the activity series and for the double replacement reactions, use the solubility rules.
    - For the reactions that occur, write the net ionic equation.

2. A student placed a 3.50g piece of magnesium, which was silver and shiny, in a crucible and heated it. She observed a bright light being produced and the magnesium turned to a white ash and had a mass of 5.80g.
    - What did the magnesium react with in this reaction? Write a balance chemical reaction.
    - Fill in the data table below based on the information above.

Mass of Magnesium before reaction	
Mass after reaction	
Mass of other element	

    - Determine the empirical formula.

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Honors Chemistry Exam Review

Essential Standard 3.1: Understand the factors affecting rate of reaction and chemical equilibrium.

1. Define the following terms:
    a. Surface area

    b. Catalyst

    c. Concentration

    d. Pressure

    e. Equilibrium

    f. Activation energy

2. The more effective collisions that occur the _____ the reaction will go.

3. What are the 3 factors that affect the number of collisions?

4. How does increasing the surface area increase the number of collisions?

5. What affect does a catalyst have on the rate of the reaction?

6. What is the difference between equal rates and equal concentrations?

7. What occurs when a reaction reaches equilibrium?

8. **$SO_3(g) + H_2O(g) \rightleftarrows H_2SO_4(l)$**
At equilibrium $[SO_3] = 0.400M$ $[H_2O] = 0.480M$ $[H_2SO_4] = 0.600M$

   a. Calculate the value of the equilibrium constant.

   b. Is the forward or reverse reaction favored?

9. **$PCl_5(s) + H_2O(g) \rightleftarrows 2HCl(g) + POCl_3(g)$**
At equilibrium at 100°C, a 2.0L flask contains: 0.075 mol of PCl_5 ,0.050 mol of H_2O, 0.750 mol of HCl, 0.500 mol of $POCl_3$

   a. Calculate the Keq for the reaction.

   b. Is the forward or reverse reaction favored?

10. Consider the following equilibrium system in a closed container:

$$Ni(s) + 4\ CO(g) \rightleftharpoons Ni(CO)_4(g) \qquad \Delta H = -161\ kJ$$

In which direction will the equilibrium shift in response to each change, and what will be the effect on the indicated quantity?

	Change	**Direction of Shift** (left, right, or *no change*)	**Effect on Quantity**	**Effect** (increase, decrease, or *no change*)
(a)	add Ni(s)		$Ni(CO)_4(g)$	
(b)	raise temperature		K	
(c)	add CO(g)		amount of Ni(s)	
(d)	remove $Ni(CO)_4(g)$		CO(g)	
(e)	decrease in volume		$Ni(CO)_4(g)$	
(f)	lower temperature		CO(g)	
(g)	remove CO(g)		K	

Constructed Response Examples.

1. *Write your answers on a separate sheet of paper.*
2. *Be sure to write your name on each page.*

1. The collision theory explains why some reactions occur and some do not and why some reactions occur faster than others.
    - List the factors that affect the rate of a reaction
    - Using the collision theory explain how they affect the rate of a reaction.
    - What are some reasons why reactions do not occur?

2. Le Chatlier's principle explains why there are changes in equilibrium reactions.
    - Explain how pressure affects the equilibrium position.
    - How does adding more reactant affect the equilibrium position?
    - Why does a catalyst not affect the equilibrium position?

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Honors Chemistry Exam Review

Essential Standard 3.2: Understand solutions and the solution process.

1. Identify the following as either an acid, a base, both.
   a. H_2SO_4 _____
   b. $Ca(OH)_2$ _____
   c. $HC_2H_3O_2$ _____
   d. NaOH _____
   e. NH_3 _____
   f. HBr _____
   g. Conducts electricity. _____
   h. Tastes sour. _____
   i. Turns litmus paper blue. _____
   j. Has a pH greater than 7. _____
   k. Turns phenolphthalein pink. _____
   l. Has a pH less than 7. _____
   m. Feels slippery. _____
   n. Reacts with metals to produce hydrogen gas. _____

2. List the strong acids.

3. Why are strong acids and bases considered "strong"?

Use the chart below to answer questions 4-6.

Table M
Common Acid–Base Indicators

Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.2–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8.2–10	colorless to pink
litmus	5.5–8.2	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

4. Which indicator(s) would be red in a solution that had a pH of 3.0?

5. Which indicator(s) would be best for identifying a basic solution?

6. Which indicator would be red at a pH of 2.0 and yellow at a pH of 4.0?

7. Calculate the pH for the following:

   a. pOH = 11.20 _____

   b. $[H^+] = 1x10^{-5}M$ _____

   c. $[OH^-] = 1x10^{-3}M$ _____

   d. Which of these are acidic?

8. Calculate the pOH for the following:

   a. pH = 1.60 _____

   b. $[H^+] = 1x10^{-10}M$ _____

   c. $[OH^-] = 1x10^{-5}M$ _____

   d. Which of these are acidic?

9. Calculate $[H^+]$ for the following:

   a. pH = 9.0 _____

   b. $[OH^-] = 1x10^{-10}M$ _____

   c. pOH = 8.0 _____

   d. Which of these are basic?

10. Calculate the [OH-] for the following:

    a. pH = 5.0 _____

    b. $[H^+] = 1x10^{-6}M$ _____

    c. pOH= 6.0 _____

    d. Which of these are acidic?

11. Calculate the molarity of a solution made by dissolving 5.60mol of HCl in 4.5L of water.

12. Calculate the molarity of a solution made by dissolving 45.0g of lithium carbonate in 300.0mL of water.

13. What volume of 6.70M sulfuric acid is needed to make 500.0mL of 3.0M sulfuric acid solution?

14. What is the concentration of sodium hydroxide, if 34.50mL of 3.0M hydrochloric acid was needed to neutralize 35.0mL of sodium hydroxide?

15. Provide an example of the following types of solutions:
    a. Solid-solid
    b. Solid-liquid
    c. Liquid-liquid
    d. Gas-liquid

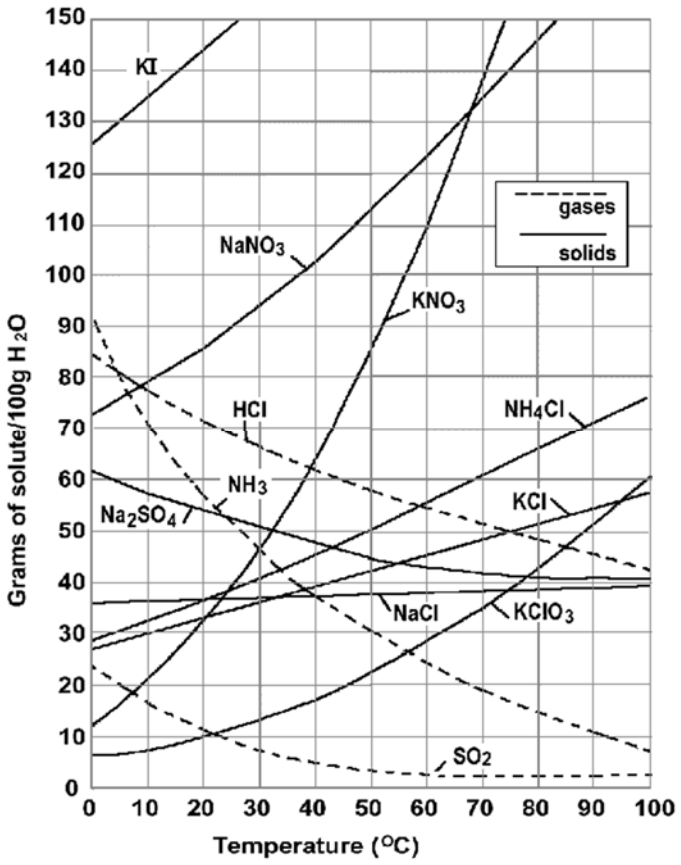
16. Define the following terms:
    a. Homogeneous
    b. Heterogeneous
    c. Electrolyte
    d. Nonelectrolyte
    e. Solute
    f. Solvent
    g. Solution
    h. Colligative property
    i. Freezing point depression
    j. Boiling point elevation
    k. Soluble
    l. Insoluble
    m. Saturated

n. Unsaturated

o. Supersaturated

17. Explain how solubility can be increased.

Use the graph below to answer questions 19-24



18. Which solid is least soluble at 10°C?

19. Which gas is most soluble at 90°C?

20. How many grams of potassium nitrate will dissolve at 50°C?

Identify the following as unsaturated, saturated or supersaturated:

21. 55g of sodium nitrate is dissolved in 100g of water at 30°C.

22. 70g of NH_3 are dissolved in 100g of water at 10°C

23. 10g of sulfur dioxide are dissolve in 100g of water at 50°C.

24. What is the relationship between the solubility of a gas and the temperature of the solution?

25. What is the relationship between the solubility of a solid and the temperature of the solution?

Constructed Response Examples.

1. *Write your answers on a separate sheet of paper.*
2. *Be sure to write your name on each page.*

1. Acids are solutions with specific characteristics
    - Describe the chemical reaction which occurs between an acid and a metal.
    - If a ribbon of magnesium and hydrochloric acid were combined, which products would result in the chemical reaction?

2. A solution containing 12.9g of $MgCl_2$ is dissolved in water to make a 0.54L solution.
    - What is the molarity of the solution? Show your work.
    - Describe how decreasing the volume would affect the molarity of the solution.