Unit 7 Review: Stoichiometry

Final Exam Review Part 2

1. What quantity is conserved in the reaction shown below? (circle all that apply)

$$2\ HCl + Ca(OH)_2 \rightarrow 2\ H_2O + CaCl_2$$

(atoms) molecules moles (mass (grams))

2. The law of conservation of mass states that in a chemical reaction ...

the mass of the reactants must equal the mass of the products

3. The quantities that ***must always*** be conserved in ALL chemical reactions are

~~# of atoms~~ and mass.

4. It takes 4 eggs and 1 cup of milk to make 2 omelets. How many eggs would it take to make 6 omelets?

$$\frac{6\ \text{omlets} \mid 4\ \text{eggs}}{1\ \text{omlet}} = 24\ \text{eggs}$$

$$3\ NH_4OH + Al(NO_3)_3 \rightarrow 3\ NH_4NO_3 + Al(OH)_3$$

5. In the reaction above, what is the mole ratio of $Al(NO_3)_3$ to NH_4OH ? 1 : 3

In stoichiometric problems, what is the mole ratio used to accomplish?

convert from moles of 1 substance to moles of another substance

$$Ca(OH)_2 + FeCl_2 \rightarrow CaCl_2 + Fe(OH)_2$$

150 g + 90 g = 128 g + ? g

(150g + 90g) - 128g = 112g

6. According to the reaction above, how many grams of $Fe(OH)_2$ should be formed? 112 g

7. Write the formula for percent yield:

$$\%\ \text{yield} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100$$

What is the amount of substance recovered from the reaction in the lab experiment called?

ACTUAL YIELD

What is the theoretical yield? Calculated amount of the product from a given amount of the limiting reactant (Maximum Yield)

SHOW ALL WORK --- BOX your ANSWER

8. $4\ Na + O_2 \rightarrow 2\ Na_2O$

a) How many moles of sodium will react completely with 3.82 moles of oxygen (O_2)?

$$\frac{3.82\text{mol } O_2 \mid 4\text{mol Na}}{\mid 1\text{mol } O_2} = 15.3\text{ mol Na}$$

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

a) How many grams of C_2H_4 (28.06 g/mol) are needed to produce 3.70 grams of water?

$$\frac{3.70\text{g } H_2O \mid 1\text{mol } H_2O \mid 1\text{mol } C_2H_4 \mid 28.06\text{g } C_2H_4}{\mid 18.02\text{g } H_2O \mid 2\text{mol } H_2O \mid 1\text{mol } C_2H_4} = 2.88\text{g } C_2H_4$$

11. $Be + 2\ HCl \rightarrow BeCl_2 + H_2$ (↓ Theoretical Yield)

Enough Be and HCl were added to produce 10.7 grams of beryllium chloride ($BeCl_2$), but only 4.50 g $BeCl_2$ was actually recovered. What is the percent yield? (↳ Actual Yield)

$$\%\text{Yield} = \frac{4.50\text{g}}{10.7\text{g}} \times 100 = 42.1\%\text{ Yield}$$

(↓ Limiting, ↙ Excess)

12. $2\ LiOH + FeCl_2 \rightarrow 2\ LiCl + Fe(OH)_2$

a) You began this reaction with 20.0 grams of lithium hydroxide (LiOH) (23.95 g/mol). What is the theoretical yield of lithium chloride (LiCl) (42.39 g/mol) ?

$$\frac{20.0\text{g LiOH} \mid 1\text{mol LiOH} \mid 2\text{mol LiCl} \mid 42.39\text{g LiCl}}{\mid 23.95\text{g LiOH} \mid 2\text{mol LiOH} \mid 1\text{mol LiCl}} = 35.4\text{g LiCl}$$

is the Theoretical Yield

b) You actually produced 6.00 grams of LiCl. What is the percent yield?

$$\%\text{Yield} = \frac{6.00\text{g}}{35.4\text{g}} \times 100 = 16.9\%\text{ Yield}$$

Review Unit 8: Energy Changes and Reaction Rates

1. Heat – energy transferred between objects due to a difference in temperatures

2. If heat energy is absorbed by a chemical system, a greater / (equal) / lesser amount of energy will be released by the surroundings.

3. In an endothermic reaction, heat is absorbed by the system.

4. In an exothermic reaction, heat is released by the system.

5. Endo solid ice melting into liquid
   Endo liquid water evaporating into gas
   Exo water vapor condensing into liquid
   Exo liquid water freezing into solid ice
   Endo solid carbon dioxide (dry ice) subliming into carbon dioxide gas

6. If you are holding a beaker in which an exothermic reaction is occurring, the beaker would feel HOT to the touch because the system is releasing energy to the surroundings which is your hand.

7. In an endothermic reaction, the reactants are at lower energy than the products.

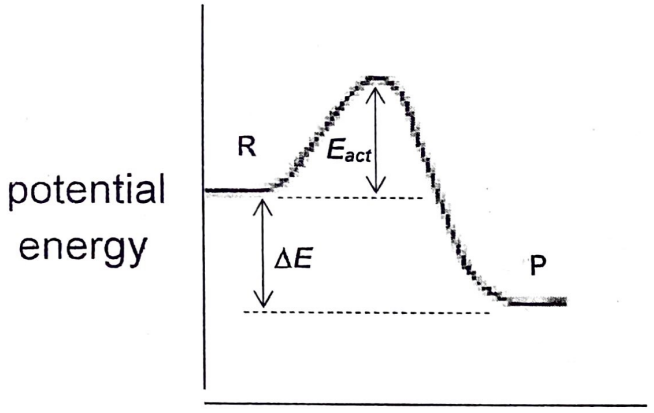
8. In an exothermic reaction, the products are at a lower energy than the reactants.

9. When chemical bonds are *formed*, energy is released;
   energy is absorbed in order to break chemical bonds.

10. In an endothermic reaction, which has stronger bonds – (reactants) or products

11. In an exothermic reaction, which has stronger bonds – reactants or (products)

12. The potential energy diagram shown is for an ENDOTHERMIC / EXOTHERMIC reaction. *(circle one)*



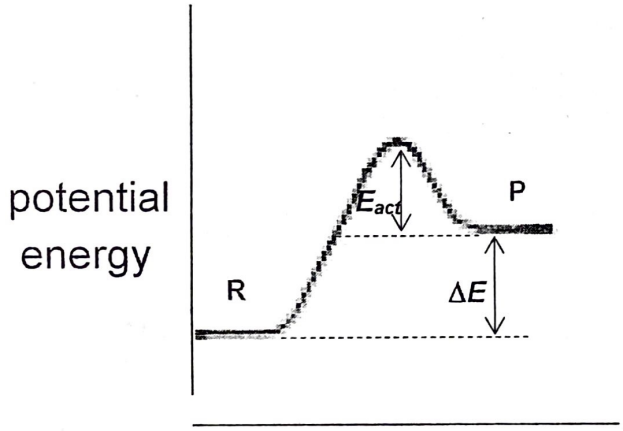
Circle the correct statement.

In an exothermic reaction, heat is released to the surroundings,

and the surroundings warm up

Touching the beaker with this reaction would feel warmer and a thermometer would show the temperature rising.

13. The potential energy diagram is for an ENDOTHERMIC / EXOTHERMIC reaction. *(circle one)*



Finish the correct statement.

In an endothermic reaction, heat is absorbed from the surroundings,

and the surroundings cool down.

Touching the beaker with this reaction would feel cooler, and a thermometer would show the temperature decreasing.

UNIT 8 REACTION RATES

1. In a typical reaction, as reaction progresses, the amount of reactants decreases, and the amount of products increases.

2. According to collision theory, a reaction can only occur if the particles collide with the proper orientation and with enough energy to react.

3. (TRUE) or FALSE (if false, change the statement to make it true)
Particles lacking the necessary kinetic energy to react bounce apart unchanged when they collide.

4. The minimum energy that particles must collide with in order to react is called _____.

A. kinetic energy
C. potential energy
(B.) activation energy
D. collision energy

5. An activated complex is an unstable arrangement of atoms that forms momentarily at the peak of the activation energy barrier. It either re-forms into the reactants or forms into products.

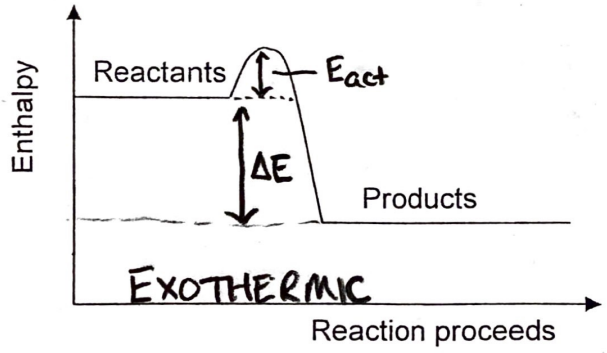
6. An increase in the temperature of a reaction will increase the rate of the reaction because more particles will collide w/ greater energy.

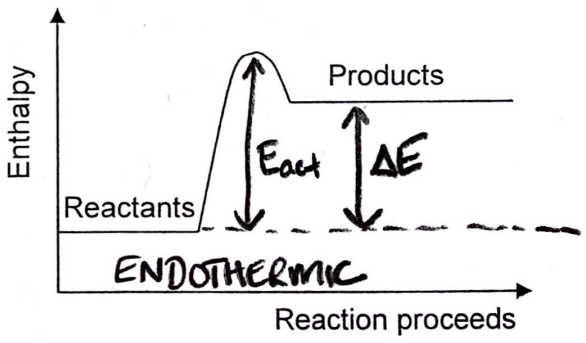
7. What happens to the rate of a reaction when some of the reactants are removed? (lower concentration) The rate decreases because there will be fewer collisions.

8. Increasing the surface area of the reactants will increase the rate of reaction. Many smaller particles have greater surface area than few large particles.

9. A substance that increases the rate of reaction without being used up itself is called a catalyst.

10. Label the Activation Energy (E_{act}) and the change in energy (ΔE)on the energy profiles below. Which is exothermic and which is endothermic?





UNIT 9 MOLECULES AND POLARITY

Review Bonding

14: Chemical Bonding Theories

Bond type	Happens between	Electrons are
Ionic	Metal & non-metal	Transferred
Covalent	Non-metals	Shared
Polar Covalent	Non-metals	Shared unevenly
Metallic	Metals	pooled

Polar covalent bond
When nonmetals bond covalently with a large difference in electronegativity

- Absolute value of differences:
  - g. 0 – 0.4 = covalent
  - h. 0.5 – 1.4 = polar covalent
  - 1.5 – 4 = ionic

Review Lewis Structures

15: Lewis Structures & VSEPR Theory

Valence Shell: Electrons in the outermost shell that bond
Octet Rule: Atoms are most stable with a full valence shell

Arranging Atoms in Lewis Structures

1. With only 2 elements, arrange symmetrically.
2. "COOH" is a carboxylic acid (both O's bond to the C and the H goes on one of the O's).
3. Hydrogen and halogens cannot go in the middle.
4. Other atoms in the order they appear in the formula.
5. Hydrogen and halogen atoms go around the element they are written next to in the formula.

Lewis Structure: A 2D representation of a molecule and its bonds.

1. Arrange the atoms as above.
2. Determine the # of valence electrons for each atom.
3. Draw the valence electrons—do not double up where a bond is going to form between two atoms.
4. Count to see if all atoms have full valences
5. If two atoms adjacent to each other do not have full valences, move in an electron from each to form a double bond. Repeat for triple bond if necessary. Move hydrogens as needed to allow double/triple bonds.

Exceptions to the Octet Rule:

1. Hydrogen and Helium can only hold 2 electrons Boron and Beryllium can be full with 6 electrons.
2. Any element in period 3 or below can have more than 8 electrons

Valence Shell Electron Pair Repulsion Theory (VSEPR): Bonds and lone pairs (electrons) repel and arrange themselves in 3D as far away from each other as possible.

The Polarity of a molecule ultimately is determined by the shape of the molecule and whether or not it is symmetrical.

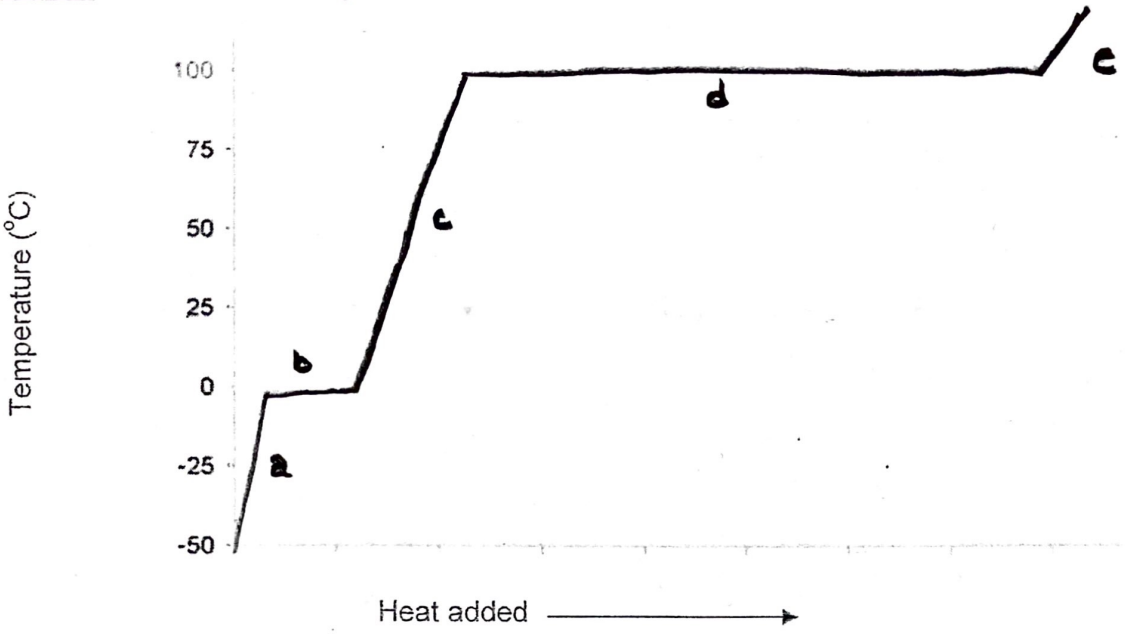
1. There are *strong attractions between polar water molecules* which cause water to have all of the following properties EXCEPT ____.
   - ✓ A. surface tension
   - ✓ B. liquid of greater density than solid (ice)
   - (C.) attraction to nonpolar molecules
   - ✓ D. higher boiling point

2. Hydrogen sulfide (H_2S) boils at –60°C. Even though water is a smaller molecule that should become a gas easier than H_2S, water doesn't boil until it reaches 100°C.

   Why do water molecules require a much higher temperature to become a gas? The stronger IMFs in water, specifically hydrogen bonds, makes much more energy required to boil water.

Unit 10 States of Matter Practice

For #23-26, refer to the heating curve below for water as heat is added at a constant rate.



23. Circle which phase(s) of water exist(s) in each section of the heating curve.

section A.	solid	liquid	gas
section B.	solid	liquid	gas
section C.	solid	liquid	gas
section D.	solid	liquid	gas
section E.	solid	liquid	gas

24. Circle which type of energy is increasing in the sample during each section as heat is being added.

section A.	kinetic	potential
section B.	kinetic	potential
section C.	kinetic	potential
section D.	kinetic	potential
section E.	kinetic	potential

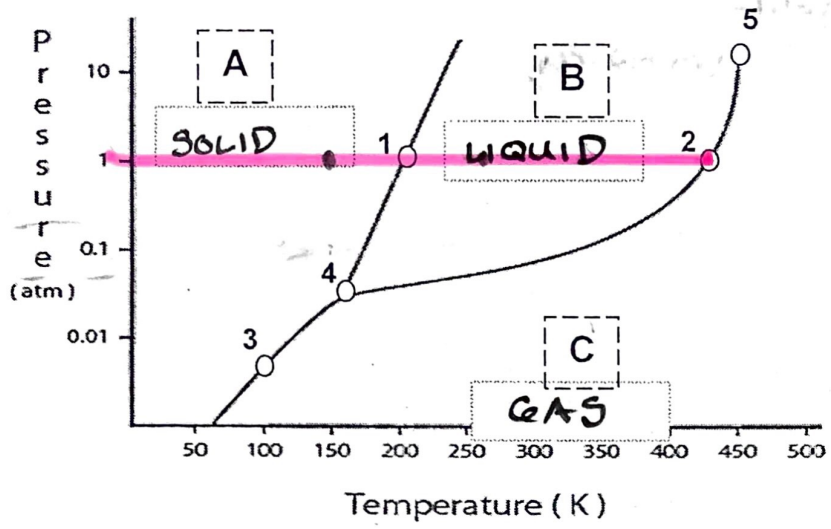
25. If heat were removed instead of added, the process occurring in section D would be _____.

A. vaporization
B. freezing
C. condensation
D. NONE of the above

26. Section B is shorter than section D because water has a lower heat of fussion than heat of vaporization so it takes more energy to vaporize a sample of water than it does to freeze the same sample of water.

27. A cup of water contains 55 g of water at a temperature of 21.4°C. How much heat must be removed from the water to lower its temperature to 2.5°C? (the specific heat of water is 4.18 J/g°C)

For #28-33, refer to the phase diagram below for water.



Label each section on the diagram (A, B, C) with the correct phase (*s*, *l*, *g*)

28. The phase change from A to C is called sublimation and from C to B is condensing

29. The boiling point of the substance is shown at Point 2 which is the point at which liquid and gas phases coexist in equilibrium.

30. Point 4 represents the triple point, which is the point at which... a substance exists in all 3 states in equillibrium

31. The critical point is shown at Point 5 which represents the temperature above which a liquid could not exist and the pressure above which a gas could not exist.

32. A sample of the substance is held constant at a temperature of 300 K while the pressure is decreased from 10 atm to 0.01 atm. The phase change that occurs is boiling.

33. A sample of the substance is held constant at a pressure of 1 atm while the temperature is increased from 150 K atm to 250 K. The phase change that occurs is melting.

Unit 11: Gas Laws

1. List 3 variables and how you would change them to increase the pressure of a gas.

(collisions with walls of container)

a) increase temperature
b) decrease volume
c) increase # of moles (molecules)

2. What happens to gas pressure if its volume is decreased? (increase) or decrease

3. What happens to the volume of a gas if the pressure is increased? increase or (decrease)

4. What happens to the volume of a gas if the temperature is increased? (increase) or decrease

5. What happens to the temperature of a gas if the volume is increased? (increase) or decrease

For #6-8, you may refer to the following relationships:

$PV = nRT$	$R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$	1 atm = 760 mmHg = 101.3 kPa
$K = {}^{o}C + 273$	STP = 273 K & 1.0 atm	Ideal gas at STP = 22.4 $L \cdot mol^{-1}$

V_1 P_1 V_2

6. 2.00 L of a gas at 2.50 atm is compressed to a volume of 0.50 L.
What is the pressure if the temperature is constant? $P_2 = ? = \frac{V_1 P_1}{V_2} = \frac{(2.00L)(2.50atm)}{(0.50L)} = 10atm$

V_1 $T_1 + 273$ P_1

7. 5.00 L of a gas at 273°C and 760 mmHg is stored in a flexible container.
What is the volume at STP? $T_2 = 273K$ $P_2 = 760mmHg$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(5.00L)(273K)}{(546K)} = 2.50\,L$$

8. A 3.50 mol sample of a gas at 305 K and a pressure of 800 mmHg. What is the volume of the gas?

n K P $\div 760 =$ 1.05atm V = ?

$$V = \frac{nRT}{P} = \frac{(3.50\,mol)\left(0.0821 \frac{L \cdot atm}{mol \cdot K}\right)(305K)}{(1.05atm)} = 83.5\,L$$

Unit 12 Solutions

Part A: Solutions & Solubility

1. A solution consists of a solute dissolved in a solvent. It is also known as a homogeneous mixture. It has a uniform composition.

2. Circle the solute and underline the solvent.
   a) 85% isopropyl alcohol in 15% water
   b) 60 mL of oil in 4 L of gasoline
   c) Sugar water

3. Solubility is a measure of the maximum amount of solute that can dissolve in a given amount of solvent at a specific temperature.

4. The 3 factors that affect how fast a solid solute dissolves (rate of dissolving) are temperature, pressure, and polarity.

5. Gases are more soluble at high pressure and low temperatures

6. What phrase describes the types of substances that will dissolve in each other: like dissolves like.
   Accordingly, polar solvents (like water) can dissolve polar solutes (alcohols, sugars, ionic compounds, etc.), and nonpolar solutes (fats, oils, hydrocarbons, etc.) will dissolve in nonpolar solvents.

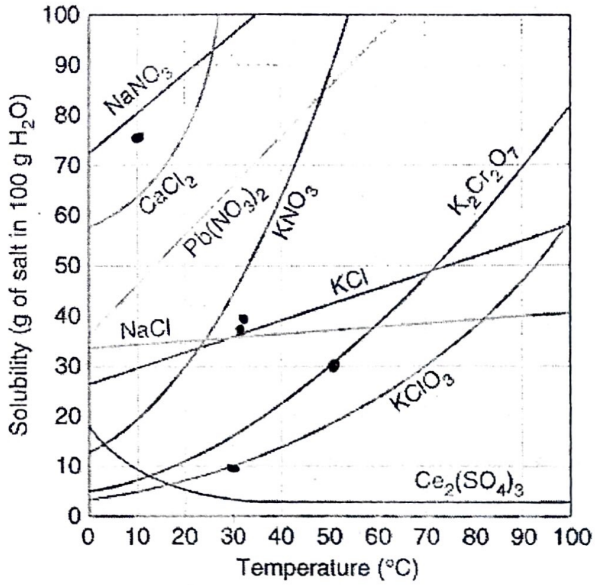
7. a) Molarity is the ratio of moles of solute per liter of solution.
   b) A 12.0 molar solution of H_2SO_4 consists of 12.0 moles of H_2SO_4 in 1 liter(s) of solution.

8. There is a large amount of solute dissolved in a concentrated / dilute solution.

9. How do you dilute a solution? Add more solvent (usually water).

10. The molarity of a solution increases / decreases / stays the same when it is diluted, and the number of moles increases / decreases / stays the same.

Use the solubility curve below to answer questions 11 - 13.



11. If ALL of the solute could be dissolved in 100 g of water at the given temperature, would the resulting solution be unsaturated, saturated, or supersaturated?

a) **40 g KCl at 30°C** supersaturated

b) **10 g $KClO_3$ at 30°C** saturated

c) **75 g $NaNO_3$ at 10°C** unsaturated

d) **35 g NaCl at 60°C** saturated

12. Which substance has the lowest solubility at 10°C? $KClO_3$

13. How many grams of $K_2Cr_2O_7$ can dissolve in 100 g of water at 50°C? 30g

Part B: Molarity & Dilutions Calculations

$$M = \frac{\text{mol}}{\text{L}} \qquad M_1V_1 = M_2V_2$$

1. What is the molarity of a solution containing 85 g NaOH in 775 mL of solution?

$$85\text{g NaOH} \left| \frac{1\text{mol NaOH}}{40\text{g NaOH}} \right. = 12.1\text{mol} \div .775\text{L} = 15.6\,M\ \text{NaOH}$$

2. How many moles are in 3.20 L of a 2.50 M solution of potassium iodide?

$$\text{mol} = L \times M = 3.20\text{L} \times 2.50\,M = 8\text{ mol KI}$$

3. How many mL of a 0.150 M_1 NaBr solution are needed to make 100 mL of 0.0500 M_2 NaBr?

$$V_1 = \frac{M_2V_2}{M_1} = \frac{(0.0500\,M)(100\text{ mL})}{(0.150\,M)} = 33.3\text{ mL}$$

UNIT 13: Acids and Bases

1. List 3 properties of acids:

taste: sour

litmus color: red

conduct (electrolytes)

List 3 properties of bases:

taste: bitter

litmus color: blue

conduct (electrolytes)

2. For the following reaction label the Bronsted-Lowry Acid, Base, Conjugate Acid and Conjugate Base:

H_2O + CO_3^{2-} → OH^- + HCO_3^-

ACID BASE CONJ. BASE CONJ. ACID

For #3 - 4, you may use the following relationships:

$pH = -\log[H^+]$	$pH + pOH = 14$	$K_w = [H^+][OH^-] = 1 \times 10^{-14}$

3. Determine the pH and Label each of the following as acidic (A), basic (B), or neutral (N).

		pH	acidic (A), basic (B), neutral (N)
a)	hydrogen ion concentration of 1 x10^{-3} *M*	3	ACID
b)	$[H^+] = 1 \times 10^{-9}$ *M*	9	BASE
c)	$[OH^-] = 1 \times 10^{-8}$ *M*	6	ACID
b)	$[H_3O^+] = 1.0 \times 10^{-7}$ *M*	7	NEUTRAL
c)	0.150 *M* hydronium ion $pH = -\log[0.150M]$	.8	ACID

4. Calculate the pOH for a solution of pH = 1.80.

$14 - pH = pOH$

$14 - 1.80 = 12.20$

For #5 - 7, you may use the following relationships:

$pH = -\log[H^+]$	$pH + pOH = 14$	$K_w = [H^+][OH^-] = 1 \times 10^{-14}$

5. Calculate the pH for a solution of 1×10^{-9} $[OH^-]$.

$$\frac{1\times10^{-14}}{1\times10^{-9}} = 1\times10^{-5} \quad \text{so} \quad pH = 5$$

6. Calculate the $[H^+]$ for a solution of 9.16×10^{-8} *M* $[OH^-]$.

$$pOH = -\log[9.16\times10^{-8}]$$
$$= 7.04$$
$$14 - 7.04 = 6.96 = pH$$

For #7, you may use the following relationship:

$$M_A V_A = M_B V_B$$

7. 10.0 mL of B NaOH of unknown concentration is titrated by adding exactly 15.8 mL of 0.150 *M* HCl to completely neutralize the base.
   A

   a. Write the balanced equation for the neutralization of this reaction.

$$NaOH + HCl \rightarrow NaCl_{(aq)} + H_2O$$

   b. What was the concentration of NaOH?

$$(15.8\,mL)(0.150\,M) = (10.0\,mL)(M_B)$$

0.237 M NaOH