Name: Ky Date: ______________ Class Pd. _____

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

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

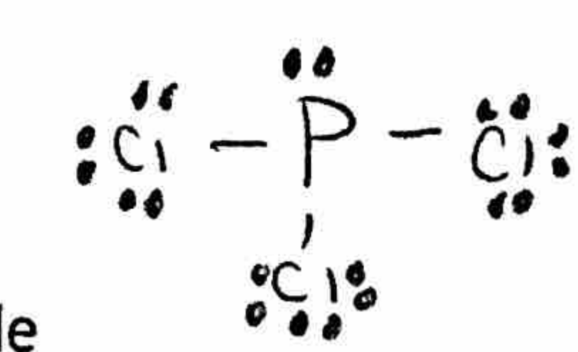
3. If an element has 6 valence electrons what charge will it likely form? −2

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

5. Why do atoms form ions? to form a stable structure / fill valence shell

6. Draw the lewis structure, identify the geometry, and identify the polarity for the following:
   a. Phosphorus trichloride PCl_3

$ve = (5) + (7 \times 3) = 26\,ve^-$

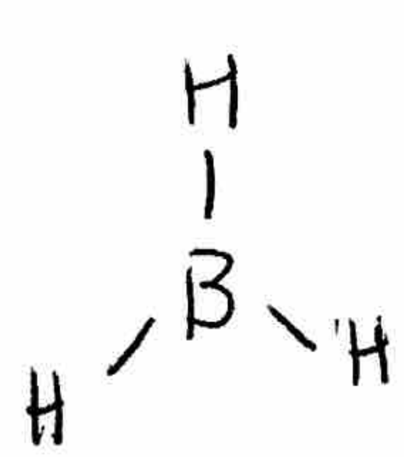


- trigonal pyramid
- polar

   b. Boron trihydride

~~Boro~~ BH_3

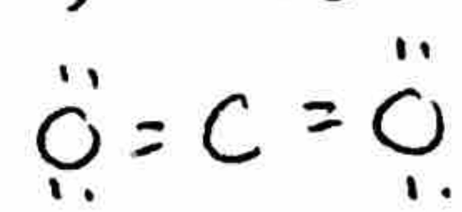
$ve = (3) + (1 \times 3) = 6$



- trigonal planar
- nonpolar

c. Carbon dioxide CO_2

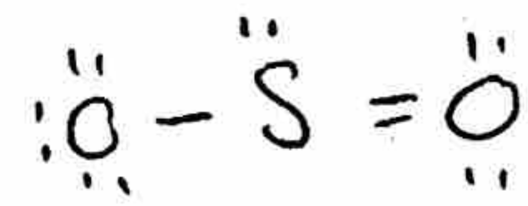
ve = $4 + (6 \times 2) = 16 ve^-$



- linear
- non polar

d. Sulfur dioxide SO_2

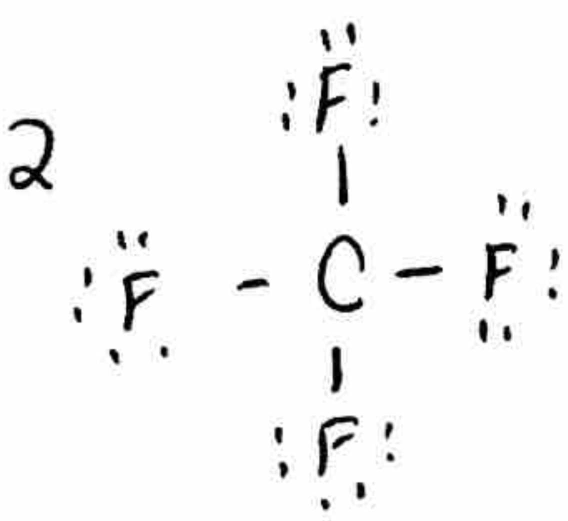
ve $= 6 + (6 \times 2) = 18 ve^-$



- bent
- polar

e. Carbon tetrafluoride CF_4

ve $= 4 + (7 \times 4) = 32$



- tetrahedral
- non polar

7. List the 7 diatomic elements.

Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

8. Write the name for the following:

a. NaBr — Sodium bromide

b. $Ca(NO_3)_2$ — calcium nitrate

c. Li_2SO_4 — Lithium sulfate

d. $FeBr_2$ — Iron (II) bromide

e. $Be(OH)_2$ — beryllium hydroxide

f. SnO_2 — tin (IV) oxide

g. N_2S — dinitrogen monosulfide

h. PH_3 — phosphorus trihydiide

i. P_2Br_4 — diphosphorus tetra bromide

j. $HClO_3$ — chloric acid

k. H_2SO_3 — sulfurous acid

l. HBr — hydro bromic acid

9. Write the formula for the following:

a. potassium iodide KI

b. magnesium acetate $Mg(C_2H_3O_2)_2$

c. aluminum chloride $AlCl_3$

d. nickel (III) nitrate $Ni(NO_3)_3$

e. calcium carbonate $CaCO_3$

f. lead (IV) sulfate $Pb(SO_4)_2$

g. beryllium phosphide Be_3P_2

h. Iron (III) carbide Fe_4C_3

i. dinitrogen trioxide N_2O_3

j. phosphorus pentafluoride PF_5

k. sulfur dibromide SBr_2

l. diboron tetrahydride B_2H_4

m. acetic acid $HC_2H_3O_2$

n. nitric acid HNO_3

o. hydroiodic acid HI

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

polar dissolves in polor
non-polar dissolves in non-polar

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

triple

12. Define the following terms:
   a. Intermolecular forces

      The forces that hold together molecules that are next to each other
   b. Hydrogen bonding

      the IMF that is created between molecules when H is bonded ~~[illegible]~~ with N, O, or F
   c. Dipole-dipole force

      the IMF that holds together polar molecules
   d. London dispersion forces

      the IMF that holds together non-polar molecules

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

    H-bonding

    London Dispeision

Constructed Response Examples.

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

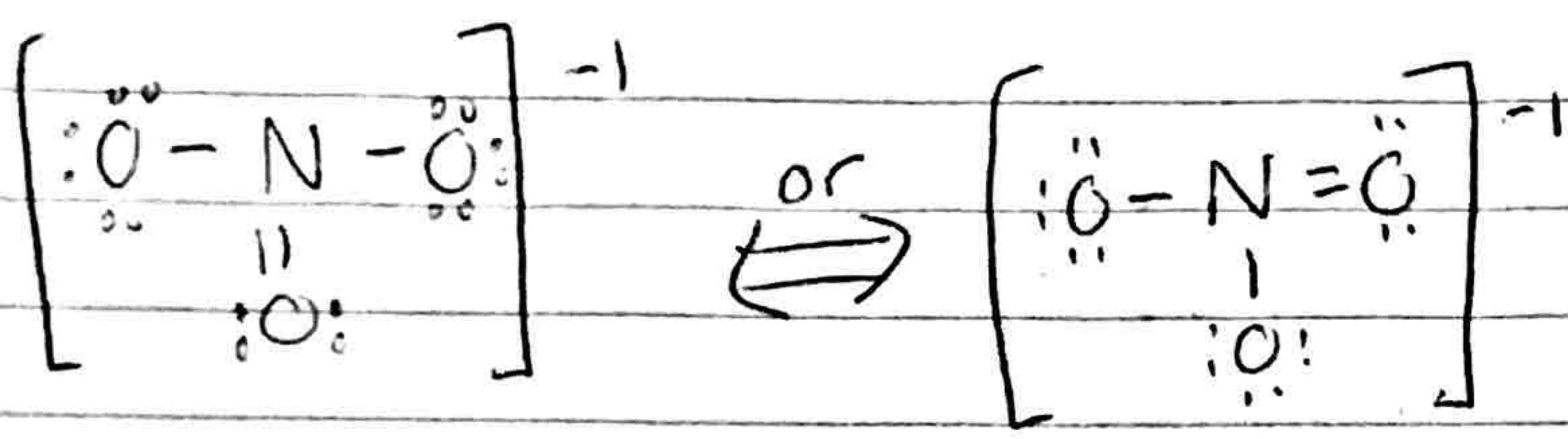
1. Polyatomic ions are ion which consist of many atoms and are common in ionic compounds.
   - Draw the lewis structures for the following polyatomic ions: nitrate, sulfate, carbonate, and ammonium
   - Identify the geometry and polarity for each of the ions above.
   - Write the formula and name for 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?

Review Material 1.2

① Nitrate NO_3^{-1}

$ve^- = (5) + (6 \times 3) + 1 = 24 ve^-$

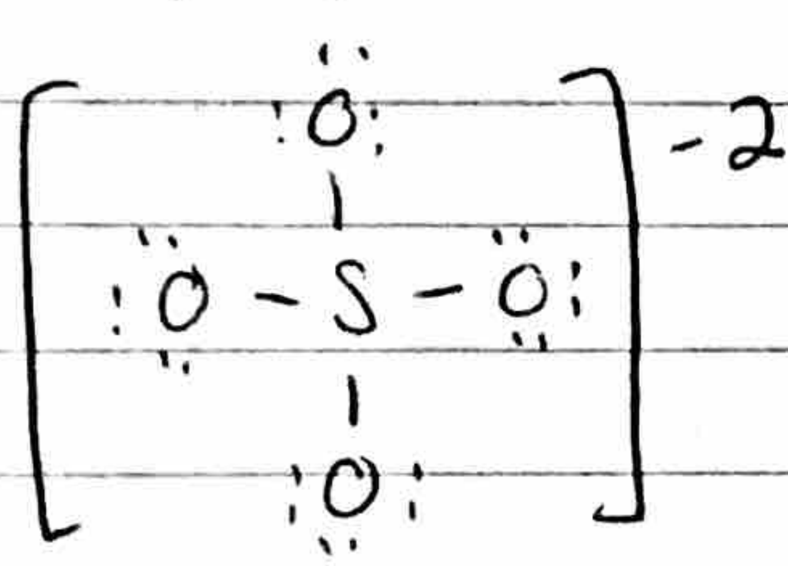


trigonal planar, non-polar, sodium nitrate

$NaNO_3$

sulfate SO_4^{-2}

$ve^- = (6) + (6 \times 4) + 2 = 32$

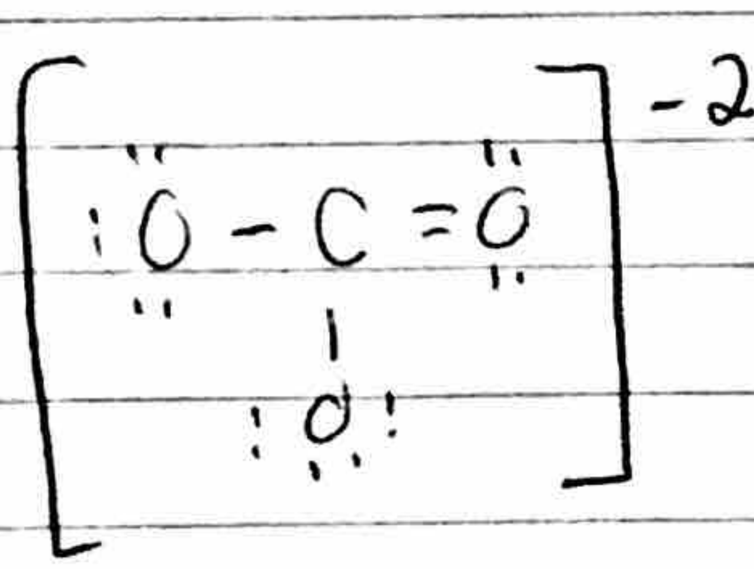


tetrahedral, non-polar, calcium sulfate

$CaSO_4$

carbonate CO_3^{-2}

$ve^- = (4) + (6 \times 3) + 2 = 24 ve^-$

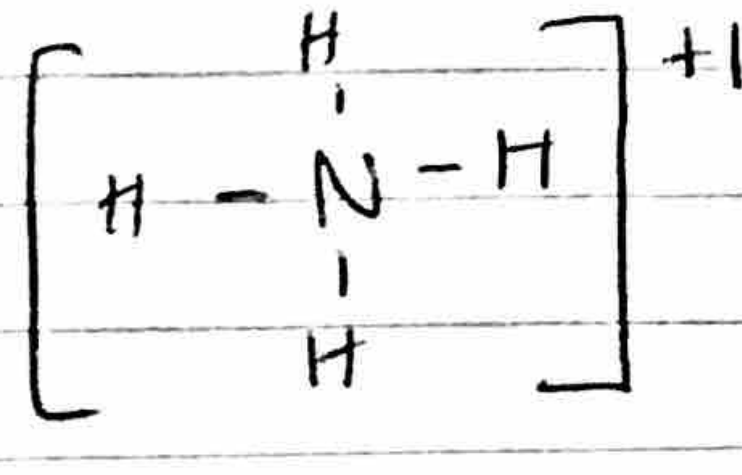


trigonal planar, non-polar, Iron(II) carbonate

$FeCO_3$

NH_4^+ ammonium

$ve^- = 5 + (1 \times 4) - 1 = 8 ve^-$



tetrahedral, non-polar, ammonium bromide

NH_4Br

(2) Ionic → transfer of electrons from cation to anion
covalent → sharing of electrons between two non-metals
metallic → sea of electrons

Ionic
→ high melting/boiling point
→ conducts electricity is aq + l state
→ solid at room temp

Covalent
→ never conducts electricity
→ low melting + boiling points
→ all states of matter at room temp

metallic
→ high melting + boiling points
→ conducts electricity in all states of matter
→ ductile + malleable

* covalent is the weakest, which is why it has the lowest melting + boiling points

* intermolecular forces are just attractions, where in intramolecular forces (bonds) there are electrons holding the compounds together

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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 periods The columns on the periodic table are called groups/families.

2. Fill in the table below:

Group	Name	Valence Electrons	Charge (oxidation state)
IA	alkali metals	1	+1
IIA	alkaline earth	2	+2
VIIA	halogens	7	-1
VIIIA	noble gases	8	N/A
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 Francium and the most active nonmetal is fluorine.

4. Metals are on the left side of the periodic table. Nonmetals are on the right side of the periodic table. Metalloids are along the staircase.

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

   a. Fe metal   b. Si metalloid   c. Ar nonmetal

   d. Ca metal   e. U metal   f. O nonmetal

6. Define the following terms:
   a. atomic radius the size of an atom
   b. Ionic radius the size of an ion
   c. Electronegativity the ability to attract an electron
   d. Ionization energy the amount of energy needed to remove an electron from the valence shell
   ~~e. Electron affinity~~

   Don't worry about this

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.

↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑

1s 2s 2p 3s 3p 4s 3d 4p

12. How many valence electrons do the following have?

a. $1s^2 2s^2 2p^6 3s^1$ 1

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

c. $1s^2 2s^2 2p^6 3s^2 3p^3$ 5

d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ 7

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

a. $1s^2 2s^2 2p^6 3s^1$ lose 1

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

c. $1s^2 2s^2 2p^6 3s^2 3p^3$ gain 3

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

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

a. $1s^2 2s^2 2p^6 3s^1$ +1

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

c. $1s^2 2s^2 2p^6 3s^2 3p^3$ -3

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

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

a. C, Fe, Fr Fr, Fe, C

b. Xe, Co, Na Na, Co, Xe

↑ very difficult to pull an electron off of a noble gas b/c it has a full energy shell

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

a. C, Fe, Fr ~~Fr~~, C, Fe, Fr

b. Xe, Co, Na Co, Na, Xe

very low electronegativity b/c Noble gases already have a full shell and don't want to gain anymore

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. high to low
   - 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?

1. → atomic size varies because elements have different # of protons, electrons and energy levels

→ Na, Mg, Al, Si, P, S

→ The smaller the atomic radius, the greater the Ionization energy

2.

metals	non-metals	mettalloids
conduct electricity	brittle	semi-conductors
malleable	do not conduct	some properties of metals and nonmetals
ductile	various colors	

→ metals become more reactive to the bottom left of PT, nonmetals are more reactive on the top right of table

→ alkali metals are more reactive than alkaline earth metals because they only need to lose 1 electron instead of two

→ halogens need one more electrons, where noble gases already have a full shell

Name: ______________________ Date: ______________________ Class Pd. ______

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	nucleus	+	about equal to n^0
Neutron	nucleus	neutral	about equal to p^+
Electron	electron cloud	-	almost nothing

2. Identify the following elements.

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

3. Write the isotopic symbol for the following:
   a. An element that has 17 protons, 18 electrons, 18 neutrons $^{35}_{17}Cl^{-1}$
   b. An element that has a 20 protons, 18 electrons, and 21 neutrons $^{41}_{20}Ca^{+2}$
   c. An element that has 93 protons, 93 electrons, and 154 neutrons $^{247}_{93}Np$

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%

$(45.699 \times .3326) + (46.799 \times .4422) + (47.899 \times .2252) = 46.68$ amu

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?
   103 nm
   b. What is the type of light emitted?
   UV
   c. Calculate the frequency of the emitted light.
   $c = \lambda \nu$
   $103\text{ nm} \left| \frac{1\text{ m}}{1 \times 10^9\text{ nm}} \right. = 1.03 \times 10^{-7}$
   $2.998 \times 10^8\text{ m/s} = (1.03 \times 10^{-7}\text{ m})(\nu)$
   $\nu = 2.91 \times 10^{15}\text{ Hz}$
   d. Calculate the energy of the emitted light.
   $E = h\nu$
   $E = (6.626 \times 10^{-34}\text{ J}\cdot\text{s})(2.91 \times 10^{15}\text{ Hz})$
   $E = 1.93 \times 10^{-18}\text{ J}$

6. Write complete electron configurations for the following:

a. O $1s^2 2s^2 2p^6 3s^2 3p^4$

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

c. Ca^{+2} $1s^2 2s^2 2p^6 3s^2 3p^6$

d. P^{-3} $1s^2 2s^2 2p^6 3s^2 3p^6$

7. Write noble gas configurations for the following:

a. Cs $[Xe]\ 6s^1$

b. Br $[Ar]\ 4s^2 3d^{10} 4p^5$

8. Define the following terms:

a. Quanta

b. Energy

c. Wavelength

d. Frequency

e. Energy level

f. Photons

look these up :)

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

10. When electrons release energy they become relaxed 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	$^{4}_{2}He$	4	+2	very low
Beta	$^{0}_{-1}e$	0	-1	low
Gamma	$^{0}_{0}\gamma$	0	0	very high

15. Complete the following reactions.

a. $^{234}_{92}U \rightarrow ^{4}_{2}He + \underline{^{230}_{90}Th}$

b. $^{18}_{8}O + ^{0}_{-1}e \rightarrow \underline{^{17}_{7}N} + ^{1}_{0}n$

c. $^{1}_{1}H + ^{2}_{1}H \rightarrow \underline{^{3}_{2}He}$

d. $^{14}_{6}C \rightarrow ^{0}_{-1}e + \underline{^{14}_{7}N}$

16. Write balanced nuclear equations for the following.

a. Alpha bombardment of Thorium-290.

$^{290}_{90}Th + ^{4}_{2}He \rightarrow ^{294}_{92}U$

b. Beta decay of Radium-222

$^{222}_{88}Ra \rightarrow ^{0}_{-1}e + ^{222}_{89}Ac$

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

$A = \frac{A_o}{2^n}$

$n = \frac{\text{total time}}{1/2 \text{ life}}$

$44.3mg = \frac{100.0mg}{2^n}$

$(2^n) \cdot \frac{44.3}{44.3} = \frac{100.0}{44.3}$

$\log 2^n = \log 2.26$

$\frac{n \cdot \log 2}{\log 2} = \frac{\log 2.26}{\log 2}$

$n = 1.18$

$1.18 = \frac{24}{x}$

$x = \frac{24}{1.18}$

$x = 20.3 \text{ hr}$

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?

$A = \frac{A_o}{2^n}$

$n = \frac{\text{total time}}{1/2 \text{ life}}$

$n = \frac{100000}{5730} = 17.5$

$A = \frac{25.0g}{2^{17.5}}$

$A = 1.35 \times 10^{-4} g$

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? → 48 hours

$A = \frac{A_o}{2^n}$

$n = \frac{\text{total time}}{1/2 \text{ life}}$

$n = \frac{55.3 \text{ hours}}{48 \text{ hours}}$

$n = 1.15$

$5.0g = \frac{x}{2^{1.15}}$

$x = 11g$

20. Identify the following as either fission or fusion.

a. Occurs in the stars like the Sun ______ fusion

b. Used to generate energy we use in our homes. ______ fission

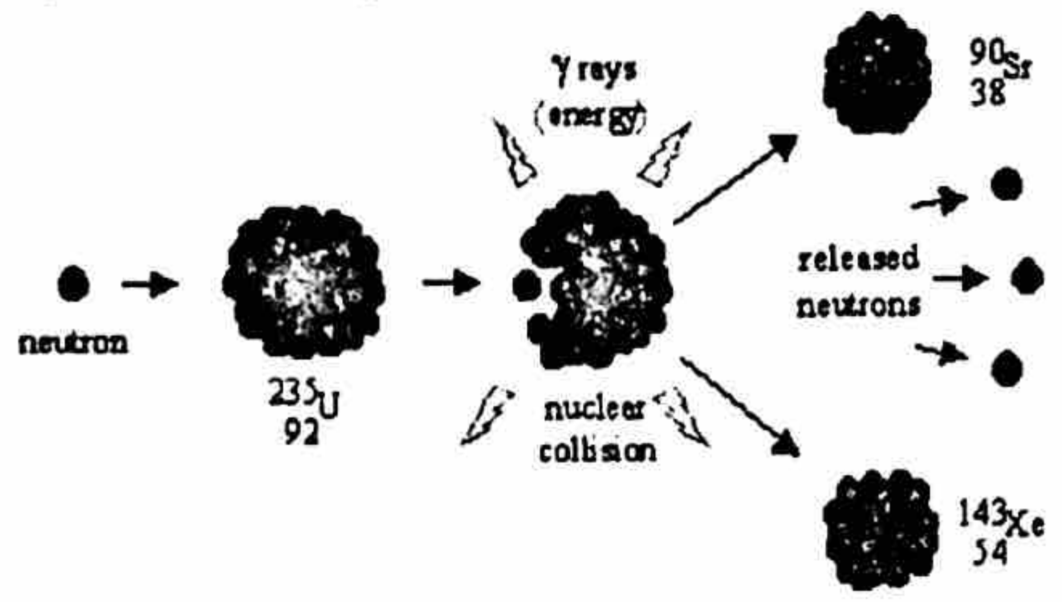
c. Combining two small nuclei to form a larger nucleus. ______ fusion

d. Creates more energy. ______ fission

e. Requires more energy. ______ fusion

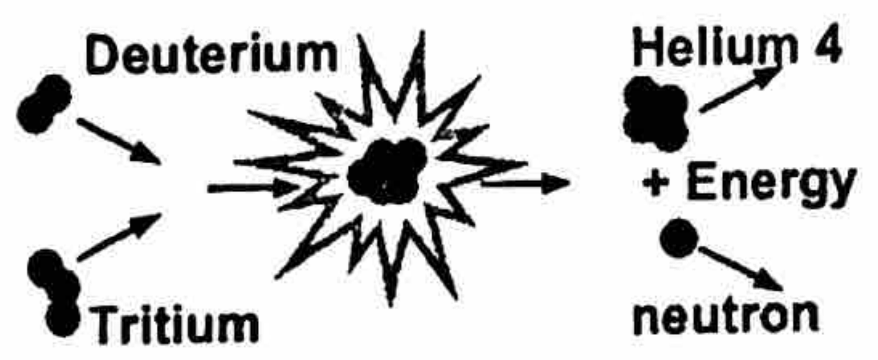
f. Splits a large nucleus into smaller nuclei. ______ fission

g.



______ fission

h.



______ fusion

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 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

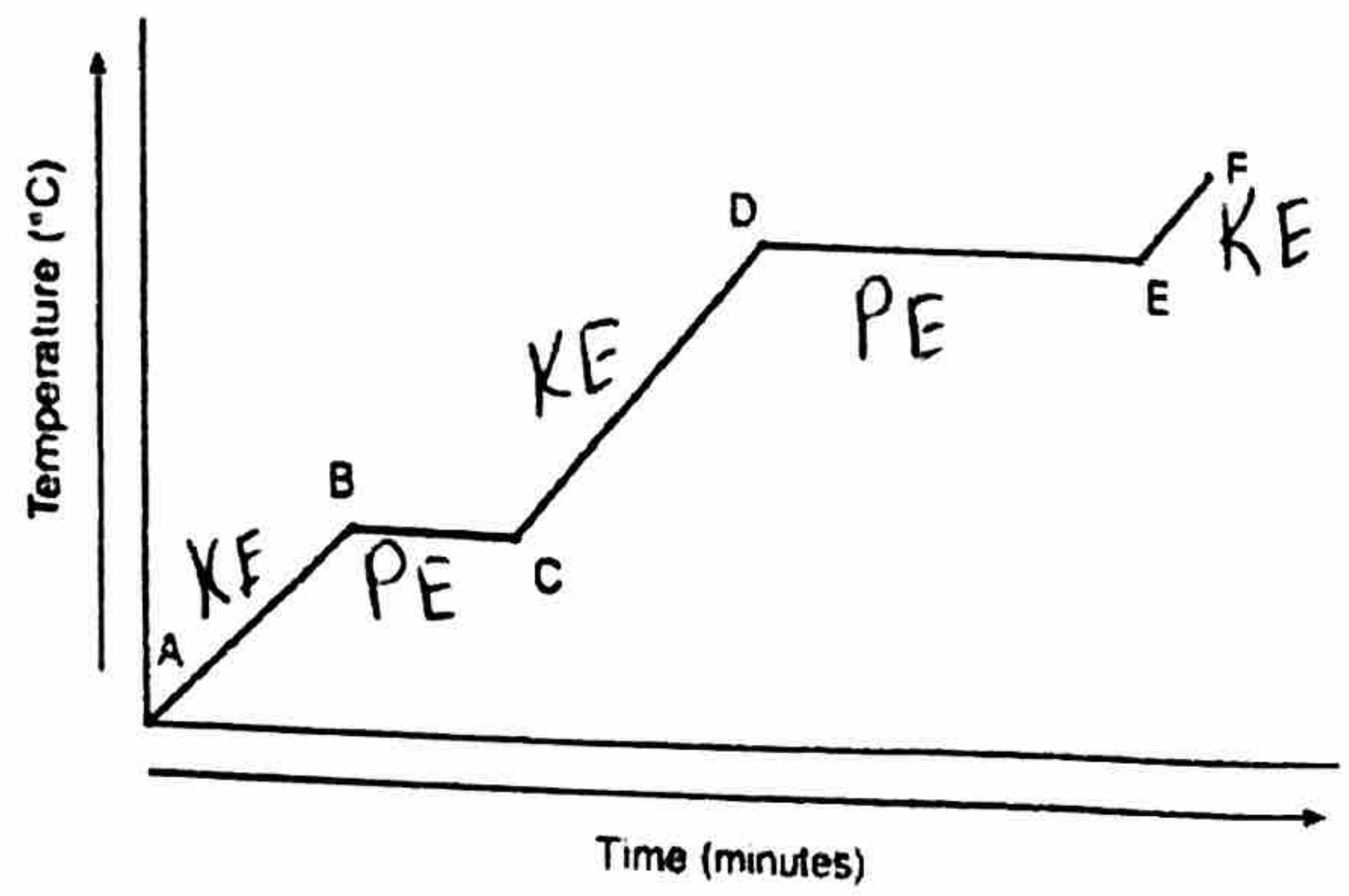
   q. Specific heat

   r. Heat of fusion

   s. Heat of vaporization

look these up in your packets

2. Identify the following on the heating curve below:



Solid: AB

melting: BC

Liquid: CD

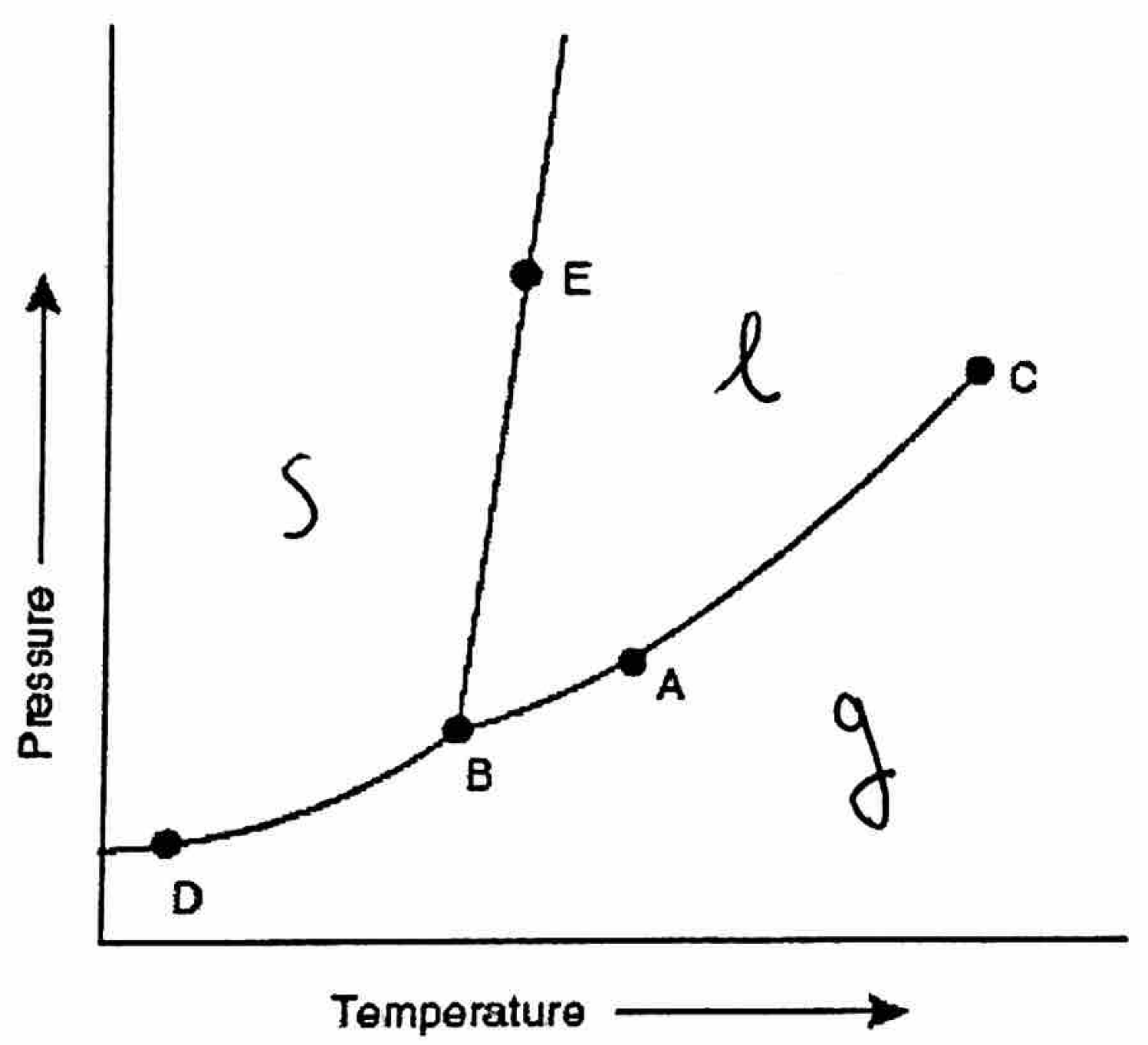
Condensation: ED

Vapor: EF

Which regions represent changes in kinetic energy? (any place where temp is changing)

Which regions represent changes in potential energy? (no temp change)

3. Identify the following on the phase diagram below



Sublimation: D~~B~~

Melting: ~~B~~E

Boiling: A

Triple point: B

Critical point: C

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? deposition

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

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

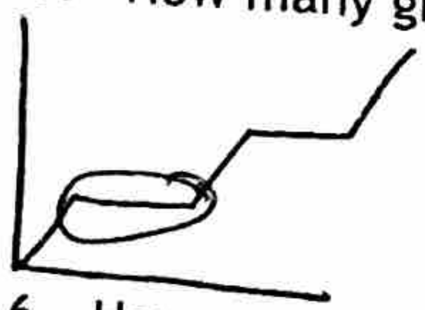
4. Calculate the mass of aluminum that would increase its temperature from 30.0°C to 40.0°C when 2500J of energy are absorbed. * use reference packet to look up specific heat of Al *

$q = mC\Delta T$

$2500\,J = m \cdot 0.897\ J/g°C\,(40.0°C - 30.0°C)$

$m = 280g$

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

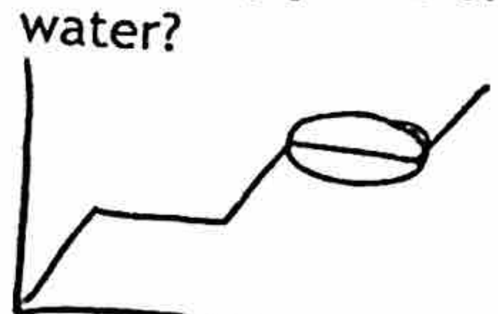


$\Delta H = mH_f$

$3500J = m \cdot 334\ J/g$

$m = 10.g$

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



$\Delta H = mH_v$

$= (150.0g)(2260\ J/g)$

$\Delta H = 3.39 \times 10^5\ J$

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

$q = mc\Delta T$

$3400J = 25.0g \cdot c \cdot (25 - 10)$

$C = 9.07\ J/g°C$

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?

know

$-q_{system} = q_{surrounding}$

9. Identify the following gas law equations:

a. PV= nRT ideal

b. $P_1V_1 = P_2V_2$ Boyles

c. $P_1/T_1 = P_2/T_2$ gay-Lussac's

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

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

f. $P_T = P_1+P_2+P_3+...$ Dalton's

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

$\frac{PV}{P} = \frac{nRT}{P}$

$V = \frac{nRT}{P}$

$V = \frac{(5.50mol)(0.0821\ \frac{L\,atm}{mol\,K})(273K)}{1.00atm}$

$V = 123\ L$

10. How many moles of argon occupy 3.40L at 1.2atm and 25.0°C? $+273 = 298\,K$

$$\frac{PV}{RT} = \frac{nRT}{RT} \qquad \frac{PV}{RT} = n \qquad \frac{(1.2atm)(3.40L)}{(0.0821\frac{L\,atm}{mol\,K})(298K)} = n$$

$n = 0.17$ mol Ar

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?

	1	2
P	1.5atm	?
V	5.6L	4.8L
T		

$P_1V_1 = P_2V_2$

$(1.5atm)(5.6L) = (x)(4.8L)$

$x = 1.8atm$

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

	1	2
P		
V	25L	?
T	1473K	298K

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{25L}{1473K} = \frac{x}{298K}$$

$$\frac{7450}{1473} = \frac{1473x}{1473}$$

$5.1L = x$

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° 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° C?

	1	2
P	1.05atm	0.65atm
V	5.0L	?
T	293K	258K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(1.05atm)(5.0L)}{293K} = \frac{(0.65atm)(x)}{258K}$$

$1354.5 = 190.45x$

$x = 7.1\,L$

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?

N_2

$P_1V_1 = P_2V_2$

$(.75atm)(5L) = x(13L)$

$x = 0.29$ atm

O_2

$P_1V_1 = P_2V_2$

$(1.25atm)(8L) = x(13L)$

$x = 0.77$ atm

$P_{total} = 0.29atm + 0.77atm$

$P_{total} = 1.06$ atm

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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 collide with enough energy and in the correct orientation.

Use the diagrams below to answer questions 2-5

Diagram A

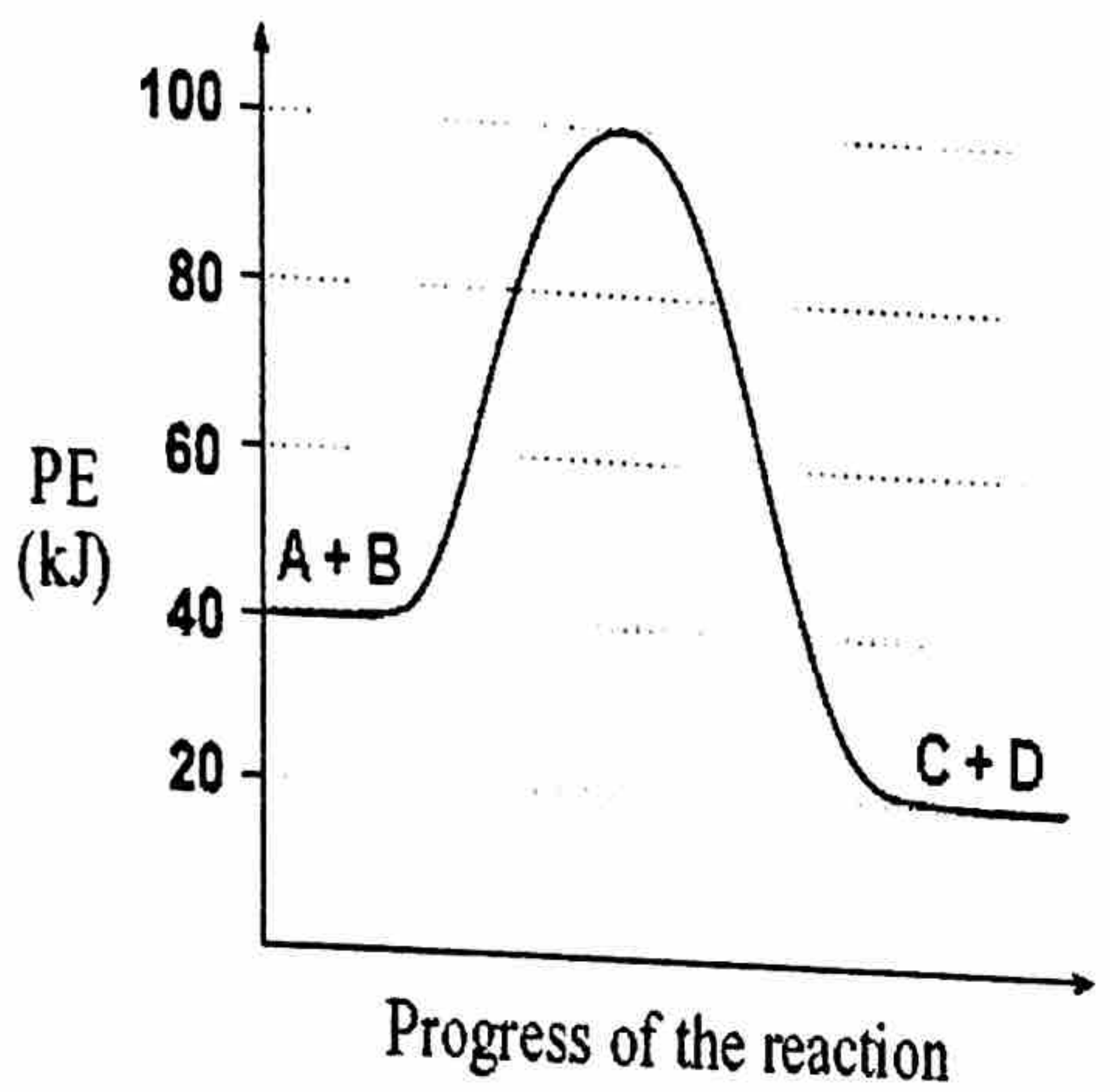
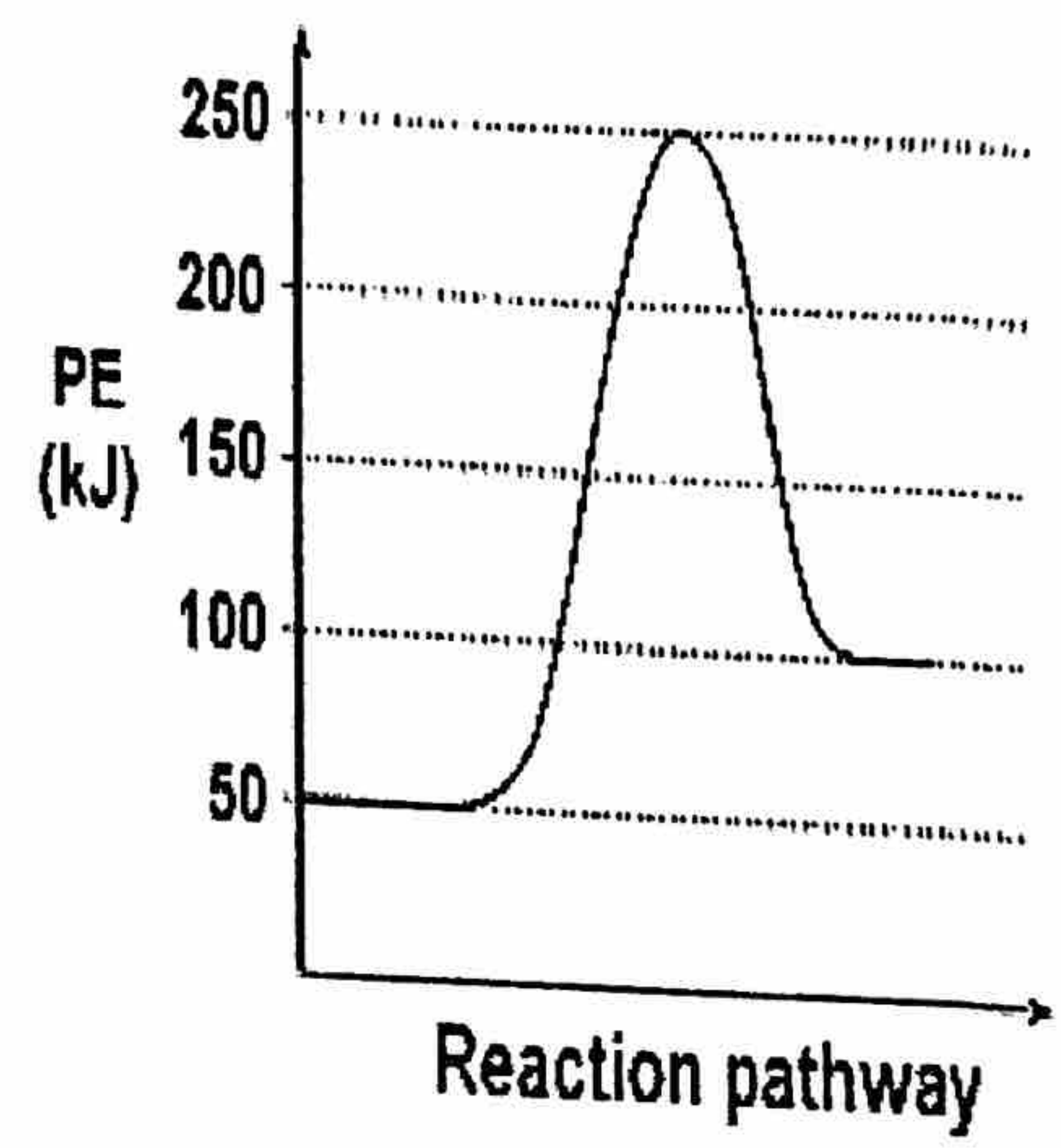


Diagram B



2. Which diagram is endothermic? B exothermic? A
3. In diagram A, what is the energy of the activated complex? 100 kJ
4. In diagram B, what is the energy of the reaction (ΔH)? 100 - 50 = 50 kJ
5. In diagram A, what is the activation energy? 100 - 40 = 60 kJ
6. The sign of ΔH is + for endothermic reactions and - for exothermic reactions.
7. List 5 indicators that a chemical reaction has occurred.

Light, heat, bubbles, precipitate, change in color

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

a. 4 Na + ___ O_2 → 2 Na_2O

Type of reaction: Synthesis

b. 2 NH_3 → N_2 + 3H_2

Type of reaction: decomposition

c. 4 C_5H_9O + 27 O_2 → 20CO_2 + 18H_2O

Type of reaction: combustion

+2 * d. 3 Pb + 2 H_3PO_4 → $Pb_3(PO_4)_2$ + 3H_2

*Pb

Type of reaction: Single replacement

e. ___ Li_3N + 3 NH_4NO_3 → 3 $LiNO_3$ + $(NH_4)_3N$

Type of reaction: double replacement

f. 3 HBr + ___ $Al(OH)_3$ → 3 HOH + $AlBr_3$

Type of reaction: double replacement (neutralization)

9. LiOH + HBr → LiBr + H_2O

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

$$\frac{10.0\text{g LiOH}}{} \left| \frac{1\text{ mol LiOH}}{23.949\text{ g LiOH}} \right| \frac{1\text{ mol LiBr}}{1\text{ mol LiOH}} \left| \frac{86.841\text{ g LiBr}}{1\text{ mol LiBr}} \right.$$

= 36.3 g LiBr

10. C_2H_4 + 3 O_2 → 2 CO_2 + 2 H_2O

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

$$\frac{4.5\times10^{22}\text{ molecules } C_2H_4}{} \left| \frac{1\text{ mols } C_2H_4}{6.022\times10^{23}\text{ molecules } C_2H_4} \right| \frac{2\text{ mol } CO_2}{1\text{ mol } C_2H_4} \left| \frac{22.4\text{ L } CO_2}{1\text{ mol } CO_2} \right.$$

= 3.3L CO_2

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?

$$\frac{5.5\text{ g NaF}}{} \left| \frac{1\text{ mol NaF}}{41.99\text{ g NaF}} \right| \frac{1\text{ mol MgF}_2}{2\text{ mol NaF}} \left| \frac{62.31\text{ g MgF}_2}{1\text{ mol MgF}_2} \right. = 4.1\text{ g MgF}_2$$

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?

$$\frac{20\text{ g HCl}}{} \left| \frac{1\text{ mol HCl}}{36.458\text{ g HCl}} \right| \frac{1\text{ mol H}_2\text{SO}_4}{2\text{ mol HCl}} \left| \frac{6.022\times10^{23}\text{ molecules H}_2\text{SO}_4}{1\text{ mol H}_2\text{SO}_4} \right.$$

$= 2 \times 10^{23}$ molecules H_2SO_4

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?

$\frac{0.0134\text{ g Fe}}{} \left| \frac{1\text{ mol Fe}}{55.85\text{ g Fe}} \right. = 0.00024 / 0.00024 = 1$

$\frac{0.00769\text{ g S}}{} \left| \frac{1\text{ mol S}}{32.07\text{ g S}} \right. = 0.00024 / 0.00024 = 1$

$\frac{0.0115\text{ g O}}{} \left| \frac{1\text{ mol O}}{16\text{ g O}} \right. = 0.00072 / 0.00024 = 3$

$FeSO_3$

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

$\frac{32.8\text{ g Cr}}{} \left| \frac{1\text{ mol Cr}}{51.99\text{ g Cr}} \right. = 0.6309 / 0.6309 = 1$

$CrCl_3$

$\frac{67.2\text{ g Cl}}{} \left| \frac{1\text{ mol Cl}}{35.45\text{ g Cl}} \right. = 1.896 / 0.6309 = 3$

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.

44.052

$\frac{88}{44.052} = 2 \times C_2OH_4$, $C_4O_2H_8$

16. What is the percent composition of potassium carbonate? K_2CO_3

K $\frac{78.2}{138.21} \times 100 = 56.6\%$

C $\frac{12.01}{138.21} \times 100 = 8.69\%$

O $\frac{48}{138.21} \times 100 = 34.7\%$

Name: ________________________________Date: ____________________Class Pd. _____

Honors Chemistry Exam Review

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

look these up !! :)

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 quicker the reaction will go.

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

   speed, energy, correct orientation

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

   more spots for effective collisions

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

   lower the activation energy

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

   - equal rates ~~the~~ means that the forward and reverse reactions are occuring at the same speed
   - equal concentrations means that there are the same concentrations of all products and reactants

7. What occurs when a reaction reaches equilibrium?

   equal rates

* l and s are not in Keq equations *

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

a. Calculate the value of the equilibrium constant.

$$K_{eq} = \frac{1}{[SO_3][H_2O]} = \frac{1}{(.400)(0.480)} = 5.21$$

b. Is the forward or reverse reaction favored?

9. $PCl_5(s) + H_2O(g) \rightarrow 2HCl(g) + POCl_3(g)$
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 $POCl_3$ → 0.025 M
→ .375 M 0.25 M

a. Calculate the Keq for the reaction.

$$K_{eq} = \frac{[HCl]^2[POCl_3]}{[H_2O]} = \frac{(.375^2)(.25)}{0.025} = 1.4$$

b. Is the forward or reverse reaction favored?

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

$Ni(s) + 4\,CO(g) \rightleftharpoons Ni(CO)_4(g) + \text{heat}$ $\Delta H = -161$ kJ ← exothermic

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)	No change	Ni(CO)4(g)	no change
(b)	raise temperature	←	K	no change
(c)	add CO(g)	→	amount of Ni(s)	↓
(d)	remove Ni(CO)4(g)	→	CO(g)	↓
(e) increase in pressure →	decrease in volume	→	Ni(CO)4(g)	↓
(f)	lower temperature	→	CO(g)	↓

Name: ____________________ Date: ____________________ Class Pd. ______

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

2. ~~List the strong acids.~~

don't worry about these :)

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

the completely ionize (break apart)

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?

methyl orange

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

bromthymol ~~base~~ blue or phenolphthalein

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

methyl orange

equations are in reference packet

7. Calculate the pH for the following:

a. pOH = 11.20 2.8

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

c. $[OH^-] = 1x10^{-3}M$ 11

d. Which of these are acidic?

a, b

8. Calculate the pOH for the following:

a. pH = 1.60 12.4

b. $[H^+] = 1x10^{-10}M$ 4

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

d. Which of these are acidic?

a

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

a. pH = 9.0 1×10^{-9} M

b. $[OH^-] = 1x10^{-10}M$ 1×10^{-4} M

c. pOH = 8.0 1×10^{-6} M

d. Which of these are basic?

a,

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

a. pH = 5.0 1×10^{-9} M

b. $[H^+] = 1x10^{-6}M$ 1×10^{-8} M

c. pOH= 6.0 1×10^{-6} M

d. Which of these are acidic?

a, b

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

$$M = \frac{mol}{L} = \frac{5.60\,mol}{4.5\,L} = 1.2\,M$$

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

$$\frac{45.0g\ Li_2CO_3}{} \left| \frac{1\ mol}{73.892g} \right. = 0.609\ mol \qquad M = \frac{0.609\,mol}{0.3000L}$$

$$\frac{300.0mL}{} \left| \frac{1\ L}{1000\ mL} \right. = 0.3000\,L \qquad M = 2.03M$$

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

$$M_1V_1 = M_2V_2$$

$$(6.70M)(x) = (3.0M)(500.0mL)$$

$$x = 220mL$$

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?

$(\#H)M_AV_A = (\#OH)M_BV_B$ $NaOH + HCl$ *1 of each*

$(1)(3.0M)(34.50mL) = (1)(x)(35.0mL)$

$3.0M = x$

15. Provide an example of the following types of solutions:

a. Solid-solid metall alloys

b. Solid-liquid salt water

c. Liquid-liquid gasoline

d. Gas-liquid soda

16. Define the following terms:

a. Homogeneous

b. Heterogeneous

c. Electrolyte

d. Nonelectrolyte

e. Solute

100k these up :)

f. Solvent

g. Solution

h. Colligative property

i. Freezing point depression

j. Boiling point elevation

k. Soluble

l. Insoluble

m. Saturated

n. Unsaturated

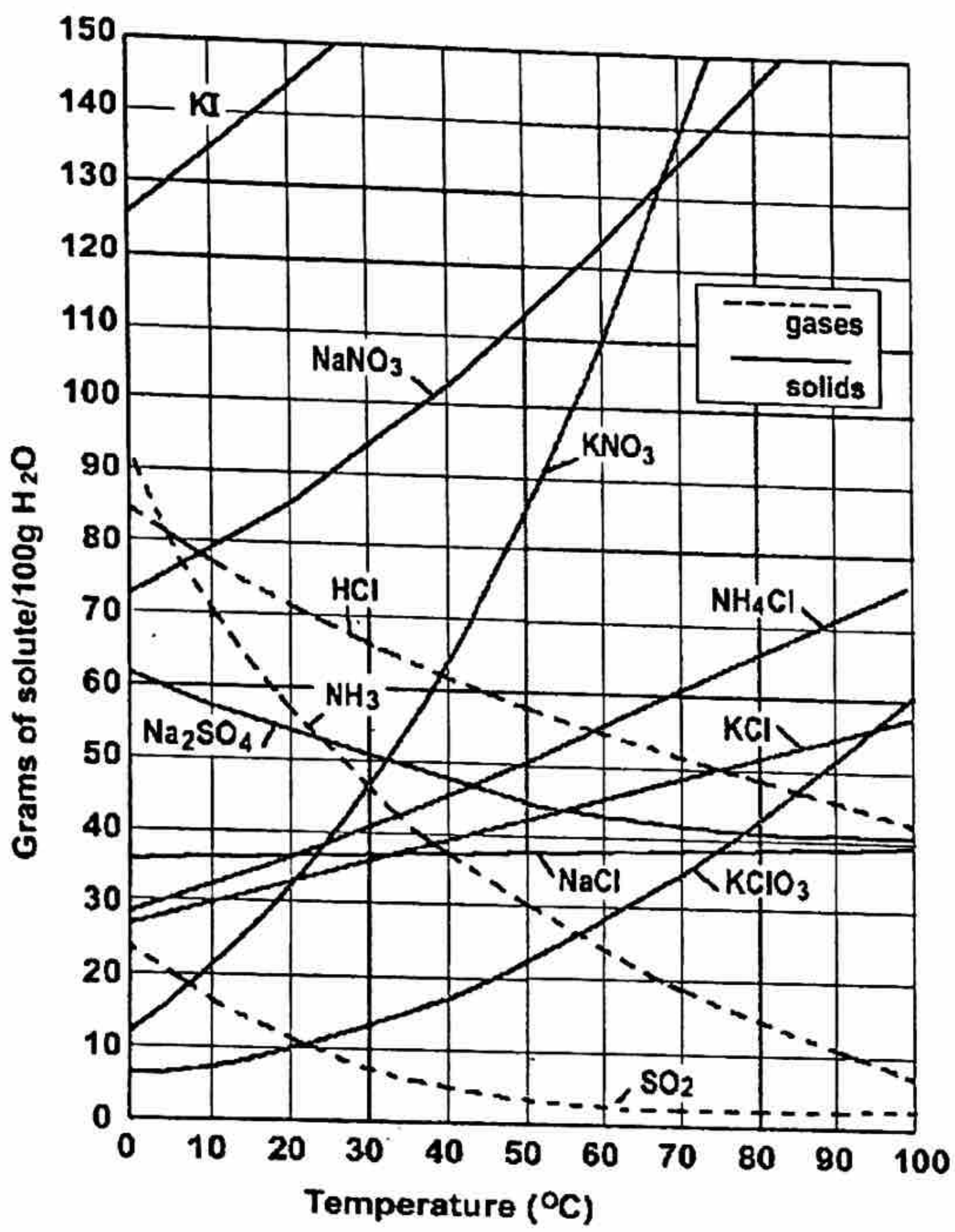
o. Supersaturated

17. Explain how solubility can be increased.

Solids → increase temp

gas → decrease temp

Use the graph below to answer questions 19-24



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

KClO3

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

HCl

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

85g

Identify the following as unsaturated, saturated or supersaturated:

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

unsatuiated

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

satuiated

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

Supersatuiated

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

inverse

→ the higher the temp, the less soluble

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

direct

→ the higher the temp, the more soluble