Name: _____

1. How much heat is gained or lost (ΔH) when 1 mole of oxygen gas (O_2) reacts in the following equation? Show the correct sign number of significant figures for ΔH. Also state whether the reaction is exothermic or endothermic.

C_2H_4 *(g)* + 3 O_2 *(g)* → 2 CO_2 *(g)* + 2 H_2O *(g)* ΔH_{rxn} = -1411 kJ/mol-rxn

ΔH =_____

Exothermic or endothermic? _____

2. Use Gibbs Free Energy (ΔG) to determine if ammonia will spontaneously boil at 0 °C. Use the data ΔH_{vap} = +23.3 kJ/mol and ΔS_{vap} = +97.1 J/mol·K. Show a calculation for ΔG.

$$\Delta G = \Delta H - T\Delta S$$

ΔG = <u>-3200 J/mol</u>

Spontaneous at 350K? <u>yes</u>

3. Draw a reaction diagram (energy vs. time) for an exothermic reaction that releases 100. kJ of energy and has an activation energy of 25 kJ. Label the reactants, products, activation energy, enthalpy change, and both axes.

4. Use the Le Chatlier principle to predict the effects on the below equilibrium. Note that all species are in the gas phase (*g*).

ΔH = -41.4 kJ/mol

a. Does the H_2O level **increase, decrease, or stay the same** when more H_2 is added?

b. What happens to H_2O when more CO is added?

c. What happens to CO_2 when H_2 is removed?

d. What happens to H_2O when more H_2 is removed?

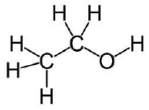
e. What happens to H_2O when the total pressure is increased?

f. What happens to H_2 when the temperature is increased?

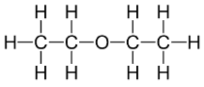
g. What happens to H_2 when a catalyst is added?

5. Indicate the strongest type of intermolecular force (IMF) for each of the following as a pure liquid.

a. water (H_2O)

b. ethanol (CH_3CH_2OH) 

c. dodecane ($C_{12}H_{26}$)

d. diethyl ether 

e. hydrogen fluoride gas HF (g)

f. ammonia (NH_3)

g. ozone (O_3)

h. nitrogen gas (N_2)

6. Which has the lower boiling point: I_2 or F_2? Why?

7. Use the concept of IMFs to explain the high boiling point of water in 1-3 sentences. You may also include a simple diagram.

8. Convert the pressure 1.012 atm to units of mm Hg. Use 1 atm = 760. mm Hg..

9. A sample of gas has a total pressure of 8086 torr and a nitrogen mole percent of 34%. Calculate the partial pressure of nitrogen.

10. The solubility of oxygen in blood 0.44 g/100 mL at sea level where the partial pressure of oxygen is 165 mm Hg. What is the solubility at a higher elevation where the partial pressure of oxygen is 65 mm Hg? Apply Henry's Law.

11. Use PV = nRT to calculate the number of moles of gas occupying a volume of 2.13 L at a pressure of 544 mm Hg and a temperature of 44.2 °C. Use R = 0.0821 L·atm/ mol·K.

12a. How much energy is gained/released when 5.32 g of water at 100 °C evaporates to form steam, given ΔH_{vap} = 40.68 kJ/mol? Show the correct sign.

ΔH =_____

12b. How much energy is gained/released when 5.32 g of water is heated from 23 °C to 100 °C? Use $\Delta H = m\ c_p\ \Delta T$ and c_p = 4.184 J/°C g.

ΔH =_____

12c. How much energy is gained/released when 5.32 g of water is both heated from 23 °C to 100 °C and evaporated?

12d. Draw a diagram of Temperature (y-axis) versus Energy (x-axis) for 12c.

13. What is the concentration in units of molarity (M) for 2.48 L of aqueous solution containing 24.6438 g of dissolved NaCl?

14. Define strong and weak electrolytes in 1-3 sentences.

15. If the pH of human blood is 7.4, calculate the hydrogen ion concentration $[H^+]$. Use $pH = -\log([H^+])$ and $[H^+] = 10^{-pH}$. Give the correct units for $[H^+]$. Watch sig figs!

16. Provide the equilibrium reaction between acetic acid (CH_3COOH) and acetate ion (CH_3COO^-) in water. Label the Lewis acid/base and conjugate base/acid.

17. Find the concentration when 175 mL of a 1.6 M LiCl solution is diluted to 1.0 L.

18. A titration experiment uses 20.55 mL of 0.300 M sodium hydroxide (NaOH) to neutralize 50.00 mL of the diprotic acid sulfuric acid (H_2SO_4). What is the concentration of the acid?

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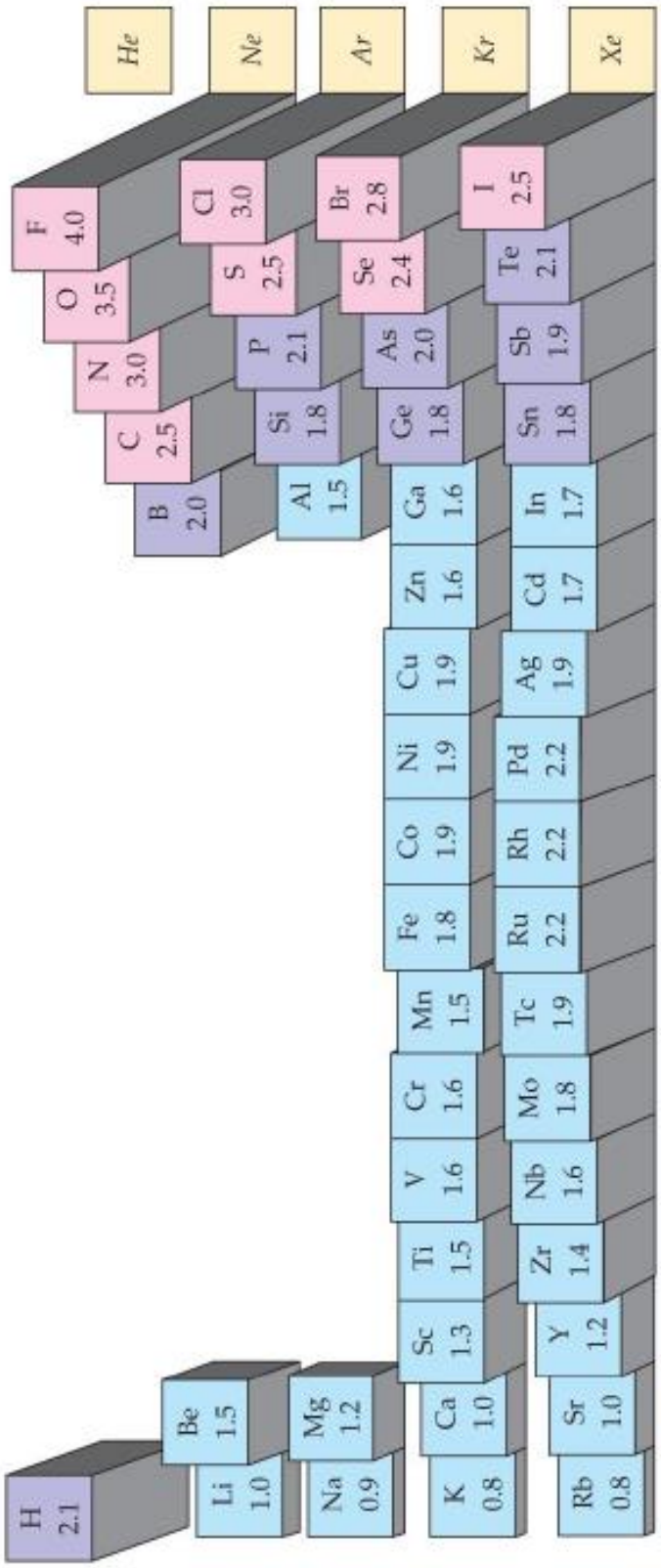

Table of Pauling electronegativity values.

Name: ___________________________

Periodic Table of the Elements

Key: 1 — atomic number; H — symbol; 1.008 — atomic weight

1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57–71*	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89–103‡	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Ds (271)								

Numbers in parentheses are atomic mass numbers of radioactive isotopes.

*lanthanide series	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
‡actinide series	89 Ac (227)	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)